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NEW ELEMENTARY CHEMISTRY

# GUIDE TO ELEMENTARY CHEMISTRY FOR BEGINNERS

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# A GUIDE

TO

# ELEMENTARY CHEMISTRY

## FOR BEGINNERS

BY

LE ROY C. COOLEY, PH. D.,

PROFESSOR OF PHYSICS AND CHEMISTRY IN VASSAR COLLEGE

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# PREFACE.

In an Elementary Chemistry, written in 1872, it was my purpose to give a short course, for beginners, in which the experimental evidence, on which the most fundamental parts of the science rested, should take the place of minute details and advanced theoretical discussions, hoping in this way to encourage the study of chemistry by experiment instead of by books alone, as was so much the custom at that day. A *Student's Guide*, printed for the use of my classes in 1878, contained a course introductory to qualitative analysis, giving the student nothing but an outline of experiments. He was expected to make the experiments, to observe and describe his own results, and from these to construct for himself a plan for the detection of the metals. I now combine the leading ideas of those two books, and offer to my fellow-teachers a new volume, in which they are more fully developed in ways suggested by the unbroken experience of the intervening years.

Chemistry as a branch of study in the schools has two great merits happily combined. One is to be found in the kind of knowledge it offers, and the other in the peculiar mental training it affords. Of these the latter is certainly not the least important, because a person is well educated, not so much in proportion to what he knows, as in proportion to what he can do with his knowledge. Hence a chief purpose

of the study of elementary chemistry in schools is to educate the mind by giving it the right kind of exercise in the use of its powers.

I have therefore tried to make a judicious selection of the most fundamental facts and principles of chemistry, and to present these in such a way that the student must constantly use his senses to discover facts, his reason in drawing correct inferences from the data he collects, and good English in expressing accurately what he sees and thinks.

I know of but one way to teach a student how to acquire a real knowledge of nature, and that is, to fix his mind habitually on things and events brought under his own eye, and direct him to the discovery of facts and principles for himself. The use of apparatus is, of course, indispensable if the student is thus to study phenomena instead of descriptions of phenomena, and the use of apparatus, by himself, is without doubt the method which is most certain to stimulate his mind to the greatest activity. Laboratory study for students in high schools is rapidly growing in favor, but unfortunately, in many schools where chemistry is taught, the difficulties in the way of this method are still thought to be real. Even in these, chemistry to be truly useful should be presented as a study of phenomena, by experiments, instead of what somebody has said about phenomena in books.

I have therefore tried to construct a course of experiments suited to the use of the beginner, at his laboratory desk, and to the use of the teacher for his class of beginners, where facilities for students to work for themselves seem to be out of reach.

The study of any subject by experiment combines two kinds of exercise; mechanical and mental operations go hand in hand. On this account experimental investigation is a com-

plex and difficult work. All that can be done to make it less so for beginners is to make one or the other, the mechanical or the mental processes, predominate in our elementary course of instruction. Then which shall it be? The mechanical of course stands first, in one sense, for there will be no phenomena to study until apparatus is selected and arranged to exhibit them. But, on the other hand, a wise selection of apparatus and conditions cannot be made by one who has not already acquired some skill in tracing the relations of cause and effect, and some experience in the application of experimental methods. I think we should first cultivate the power to observe exhaustively and to detect relations, — that we should make the mental more prominent than the mechanical in the elementary study of chemistry. Accordingly:

In this course of experiments the mechanical operations are described in quite minute details. Exactly what is to be done is told, but what is to happen, and the meaning of it, is for a time withheld. Exceptions to this plan will be found in the description of processes which are simply means to secure conditions, and in the statement of facts which may be needed for immediate use. But in general the phenomena which hold the chemistry of substances or processes are left for the student to discover. See, for example, page 35, or pages 85, 86.

I know that much stress is, by many, laid upon the industrial value of an instrument-making course in chemistry. But it seems to me that the study of chemistry is not primarily to teach mechanics, and that the use of tools and the possession of mechanical ingenuity can be better acquired in the industrial school or workshop, where these are the specific aims, than in the laboratory of the high school or academy, where the acquisition of knowledge for the sake of mental training is

the chief purpose. Home-made apparatus is not to be despised, but to be greatly respected, where nothing better can be had, for much can be done with the most common utensils, such as bottles, fruit-jars, tea-saucers, and oyster-cans. But certainly beginners can do better work with good facilities than with poor ones. And while there is so much in the market which is at once scientific and inexpensive, the student should be taught to reach more accurate results than are otherwise possible by the use of it. Productive ingenuity and skill must be founded on exact knowledge and clear thinking; they cannot precede these. Therefore:

The apparatus called for in this course has been selected from that which is made for, and approved by, chemists. The pieces are neat, simple, easily put together, always in market, and as cheap as possible for good scientific work. (See Appendix, Fig. 69.)

A brief summary of the most important facts and principles follows the experimental work, by which the student can check and correct his results. In this summary will be found the information which should be acquired by beginners in chemistry. I have tried to include in it only such things as will be of most value to the many who will finish the study of chemistry in the high school, and to the few also who are there to lay a foundation for college work. "Not how much we know is the best question, but how we have got what we know, and what we can do with it, and, above all, what it has made of us." — J. P. LESLIE.

It is not well to undertake too much. It is not best to have the student's text-book burdened with matter which he is not expected to master. There is more education to be gained by extending the search for facts into other volumes than by skipping parts of the book in use. I have not given a long

list of experiments, but have tried to make a judicious selection, believing that a few typical ones well made and thoroughly studied, are far more useful than a larger number would be if studied in haste. What I mean by the thorough study of a few experiments in the treatment of a subject may be seen by referring to "Substitution," pp. 19–21: "Decomposition of nitric acid," pp. 92–95; or "Chlorides," pp. 141–145.

Additional work is better when provided by teachers for such pupils or classes as have time or talent to undertake it. I would make such work partake of the nature of research as much as possible. A student may be given some question to be answered by his own experiments, or two substances whose mutual reactions and results he is directed to investigate, or a single body whose properties he is asked to study and report. Some work of this kind I have given under the head of "Exercises." (See, for examples, pp. 39, 82, 100.)

Next in value to research in the laboratory stands research in the library. By all means teach the student how to make the results of his study, with apparatus and the text-book, the nucleus around which to group other facts, a center from which to extend his knowledge. From the following works the teacher can select abundant materials for this exercise, in kind and quantity suited to the varying wants of different individuals or of successive classes. Buckley's "Short History of Natural Science." Wurtz' "History of Chemical Theory." Wurtz' "Atomic Theory." Cooley's "New Text-Book of Chemistry." Cooke's "New Chemistry." Remsen's "Organic Chemistry." Remsen's "Theoretical Chemistry." Roscoe and Schorlemmer's "Treatise on Chemistry." Fresenius' "Qualitative Analysis." Douglas and Prescott's "Qualitative Analysis."

I have in all cases rejected dangerous experiments, but I have in many cases devised simple, safe, and efficient ways to study explosive and noxious substances. See, for examples, Hydrogen, pp. 29, 30, and Chlorine, pp. 138, 139.

The wood-cuts which represent the experiments are, with a single exception, Fig. 23, made from the photographs c drawings of the apparatus in actual use. For the selected cuts, which illustrate the descriptions of historical or industrial work, I am unable to give the credit which is due to their unknown authors.

L. C. C.

POUGHKEEPSIE, June, 1886.

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# ELEMENTARY CHEMISTRY.

## OBSERVATION AND EXPERIMENT.

In the study of Chemistry we are to learn some things about the different kinds of matter. There are two ways in which these things have been found out, and in these same ways we can most easily learn what these things are. These two ways of studying nature are called *observation* and *experiment.*

**Observation.** — When I look at something which is going on, and watch carefully to see what happens, my act is an observation. To look at an object so closely that we can see its shape, its color, and whatever else is visible about it, is an act of observation.

If, for example, I desire to know as much as possible about a butterfly, the best way to learn it is to catch the butterfly, look at it intently, note down and remember what I see. The butterfly would show me that it has four wings, six legs, two long hair-like bodies (antennæ) reaching forward from its head with knobs upon their ends, two large, dark, and prominent eyes which do not close nor turn, and that the beautiful colors of its wings are due to a fine dust which is easily rubbed off by my fingers. All these facts I could learn by holding the insect in the hand and looking at it thoughtfully.

Knowledge which I get in this way is learned by observation.

**Experiment.** — But if, instead of only looking at an object as I find it, I do something to it to see how it will behave or appear in different conditions, this operation is an *experiment.*

Will 5 cubic centimeters of water dissolve as much as 10 grams of granulated sugar? I cannot find out by simply looking at sugar and water. In order to learn what the fact is, I may put the two things together in the right way, and if I do so I make an experiment. Thus:

*Ex. 1.* — I take a tall glass cylinder, *a*, Fig. 1, which is graduated to measure cubic centimeters, and pour in water

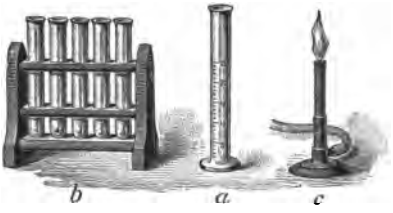



Fig. 1.

up to the 5 cc. mark.[1] I transfer this water to one of the thin round-bottomed cylinders, *b*, called a test-tube. I also weigh out 10 g. of granulated sugar[2] and put it into the water in the tube *b*.[3] I now warm the tube in the flame of a Bunsen lamp, *c*. There is danger of breaking the tube if I heat it too suddenly, or too long in one spot,

[1] If one must get along without a graduated cylinder, he may obtain 5 cc. very nearly by filling his test-tube one inch above the rounded bottom. The tube is supposed to be $\frac{5}{8}$ inch in diameter.

[2] If one must get along without a balance, he can obtain about 10 g. of dry sugar by filling a teaspoon twice.

[3] Fold a narrow strip of paper into the shape of a trough and lay this in the tube, which should be held in a slanting position. The dry sugar will slide safely down this trough instead of clinging to the wet walls of the tube.

and to avoid this danger I move it slowly in the flame to heat all sides evenly. When the liquid begins to boil I lift the tube into the hot air above the flame, where I can keep it hot without boiling it too vigorously. I watch to see

Whether the sugar remains, or becomes less and less.

Whether it all finally disappears.

If the liquid at length becomes, as it will, almost or quite transparent, we shall know that 5 cc. of *hot* water can dissolve 10 g. of sugar. I will then stand the tube in the tube-rack, and when it is cold I will look again and see

Whether 5 cc. of *cold* water can hold the 10 g. in solution.

Let us keep this syrup for use in another experiment.

The sap of some trees and the juices of some plants are *natural* solutions of sugar in water, but the quantity of sugar in 5 cc. of these juices is very small. Nothing but an experiment could have first shown that 5 cc. of water can dissolve so much sugar as we have found it to do.

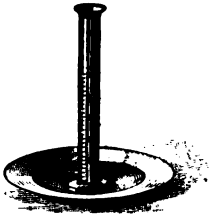


Fig. 2.

But in experiments we often put things together in ways in which nature never does. For example, I wish to know how sugar will behave in strong sulphuric acid. Nature never puts these two things together, and the only way I can find out how they will act in the presence of each other is to bring them together. Thus:

*Ex. 2.* — I measure out 5 cc. of strong sulphuric acid with the cylinder *a*, Fig. 1, pour it into an empty test-tube, then rinse the cylinder and stand it on a small plate, Fig. 2. I now pour the sugar syrup made in the other experiment into this cylinder. I am ready now to bring the two together. I pour the acid in a slender stream into the syrup, and watch for every change that happens. I notice

A change in color.

A change in volume (size).

A change in temperature (warmer or colder).

A new substance unlike sugar or syrup or acid.

As soon as the experiment is over I write, in my note-book, a short account of what I did, and the results just as I saw them.

The fact is that a coal-black, bulky mass of hot carbon or charcoal is the result of bringing these two substances together.

The science of Chemistry is founded on facts which have been discovered by experiment, and the most natural way to study Chemistry is by the same means. The best way for the student is to make the experiments himself. The second best way is to see them made by a teacher. In either case the student should remember that the object of making experiments is to discover truth. An experiment may be pretty and interesting, but its value does not lie in its beauty. No experiment is good for anything in the study of Chemistry unless it helps to reveal some truth.

The student should remember, also, that it is not what he reads about experiments, or hears a teacher say about them, that is going to give him the best and quickest insight into Chemistry, but that which he sees with his own eyes and describes in his own words.

To study Chemistry by experiment the student should obey the following rules: —

1. Arrange the apparatus and use it *exactly* as directed.
2. Watch carefully to see *every* change which takes place.
3. Note accurately on paper every *important* change.
4. Compare these results with the facts stated in the book, and correct those which are found to be wrong.
5. Study carefully to see how certain inferences may be made from the results.

## CHEMICAL CHANGES.

We already know by observation that there are changes all the time going on in bodies of matter. Some things change very rapidly, others very slowly. Wood changes to smoke and ash sometimes in a few minutes; a stone crumbles to powder only after many years. But nothing can forever stay exactly as it is.

The first thing we have to do in Chemistry is to become acquainted with these changes. How do they differ? How are they brought about, and what terms are used to describe them?

*Ex. 3.* — I take a piece of magnesium wire or ribbon about six inches long, grasp one end with a pair of pincers,[1] and hold the other end for a moment in the flame of the Bunsen lamp, Fig. 3. I see that

The metal becomes red hot, then bursts into flame.

Nothing finally remains but a crumbling white solid.

*Ex. 4.* — I now in the same way hold a piece of iron wire in the flame of the Bunsen lamp, and see that

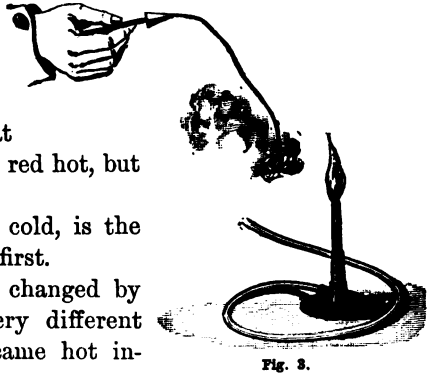


Fig. 3.

The metal becomes red hot, but does not burn.

And finally, when cold, is the same substance as at first.

Both metals were changed by the heat, but in very different ways. The iron became hot in-

[1] A homely handle can be made by starting a split in one end of a stick. The wire can be inserted in the split and held as with pincers.

stead of cold, red instead of black, but it remained iron. The magnesium did not remain magnesium, but was changed into a new substance; the crumbling white solid is a very different thing from the tough, gray, and lustrous metal. There was no change in the substance of the iron; there was a change in the substance of the magnesium.

**Observations.** — I watch the clouds and see them change in size and shape, and oftentimes in color. But their substance does not change, for clouds are made of water-vapor, whatever may be their sizes, shapes, or colors.

To-day a piece of iron is smooth and bright, but if left in moist air it will in time be found covered with iron-rust. A part of the iron becomes changed into this very different kind of matter.

Further observations show that there are changes going on all the time around us which do not alter the nature of substances, like the changes in the iron wire when heated. All such changes are called *physical changes.* There are others in which the substance of bodies is changed into matter of a different kind, like the burning of the magnesium. All such changes are called *chemical changes.* And so we learn that the multitudes of changes in the world of matter may be grouped in two great classes.



Fig. 4.

If now we go on to compare one chemical change with another we shall find that there are several different varieties.

**Decomposition.** *Ex. 5.* — I put one gram of "red oxide of mercury" into a side-neck ignition-tube,[1] Fig. 4, — just

[1] Ignition-tubes are made of "hard" glass while other tubes are made of "soft" glass. Hard glass will stand a strong heat, while soft glass will not. Common test-tubes may be used for heating liquids; ignition-tubes should be used for heating dry solids.

about enough to fill the rounded bottom. I close the mouth of the tube with a nicely fitting cork, and slip one end of a piece of rubber tubing, *e*, Fig. 5, over the end of the side-neck. I then fix the tube very obliquely in the clamp of the support, *f*. I next put a half-inch of water into the test-tube *b*, and slip the free end of the rubber tube down into it.



Fig. 5.

The oxide is now ready to be heated, and if anything shall be driven out we may catch it in the test-tube.

I now apply the flame of the lamp, and move it slowly to heat the bottom of the tube evenly. The upper part of the tube must not be heated. Now look for and describe

A change in the color of the oxide.

A coating which forms on the cold walls of the tube.

Bubbles which escape from the rubber pipe in the water.

A change in the quantity of the oxide.

Were the bubbles which came over through the water anything more than air? To answer this question I take a splinter of wood, very slender, and long enough to reach to the bottom of the tube, set fire to one end and watch the spark very closely while I push it to the bottom of the tube to see

Whether it burns just as it does in air.

THE FACTS. — By heating the red oxide of mercury it is first blackened and then broken into two kinds of matter quite unlike itself. One of these appears in shining droplets on the cold walls of the tube, the other goes off as a

colorless gas which brightens the burning of a splinter. The shining droplets which coat the cold walls of the tube are *mercury,* and the gas in which a splinter burns with unusual brightness is *oxygen.*

This is a fine example of chemical change. But the most important thing to notice is, that in this change *one substance is broken into two* which are entirely unlike itself and unlike each other. Such a chemical change is called *decomposition.*

**Decomposition of Potassium Chlorate.** *Ex. 6.* — Potassium chlorate is a white solid. Before I heat it the coarse grains or crystals should be reduced to powder: I grind it in a mortar (Fig. 6). I put two grams of the powder into the ignition-tube,[1] Fig. 4. This quantity will fill about one inch of the tube. I put three or four cubic centimeters of blue litmus solution into one test-tube, *b,* and as much lime-water into a second tube, *c,* Fig. 5, and provide a good cork for each. I put the end of the rubber tube into the litmus, and then heat the chlorate just as I did the red oxide before. Watch for and describe



Fig. 6.

The changes in the chlorate.

The bubbles from the pipe in the litmus.

After a while I put a match-flame into the mouth of the tube and see that it burns with unusual brightness. This shows that the tube is filled with oxygen.

I then put the end of the rubber tube over into the lime-water in *c,* and close *b* with its cork, in order to keep its oxygen for use further on.

At length the boiling chlorate thickens, and soon after

[1] The tube must be clean and dry. A piece of dry cloth, or a sponge tied on the end of a wire or stick, is convenient for wiping tubes.

dries up completely to a white solid. The work is done. I stop the heat, remove the rubber tube, and cork the test-tube *c*, to keep its oxygen also for future use.

Has any change been made in the litmus or the lime-water?

THE FACTS. — By heating potassium chlorate it is first melted and afterward broken into two substances, unlike itself and each other. One is the white solid, left in the ignition-tube, and the other is *oxygen*. The chlorate is *decomposed*.

Oxygen, which has appeared in both these experiments, is an important substance, and as soon as we are through with the special study of chemical changes we will examine it fully. At present we will make two experiments with it.

**Combination.** — What will happen if a bit of coal is heated in oxygen?

*Ex.* 7. — I wind the end of a small wire around a little splinter of charcoal, heat the charcoal until it holds a spark of fire, and then lower it into the oxygen in the test-tube *c*, Fig. 5, over lime-water. Notice

What effect is produced on the spark.

Whether the charcoal wastes away.

Will the oxygen brighten a match-flame afterwards?

I now cover the mouth of the tube with my finger and shake it briskly.

What change takes place in the lime-water?

THE FACTS. — Charcoal with a dull red spark will glow brightly in oxygen and burn away rapidly; the oxygen is used up at the same time, and the lime-water afterward becomes turbid and white.

But we know that oxygen will not whiten lime-water: this was proved in Ex. 6. (How was it proved?) Charcoal also will not whiten lime-water. But the burning of the

charcoal in oxygen yields something which will. It is a colorless gas called *carbon dioxide.* In place of charcoal and oxygen we have carbon dioxide, which is very unlike both.

Now the thing most important to see in this case is that *two substances are changed into one.*

*Ex. 8.* — I take a piece of wire, — perhaps a knitting-needle, — warm it a little, and then plunge it into some flowers of sulphur. A thin layer of sulphur will cling to the wire. I set fire to this sulphur, and then at once thrust it into the test-tube *b*, containing oxygen and litmus, left from Ex. 6. Notice and describe

The fine color of the flame.

The vapor which is produced.

I next close the tube with my finger and shake it to mix the vapor with the litmus.

What change is made in the color of the litmus?

Sulphur burns more freely in oxygen than in air, and with a rich blue flame, filling the vessel with white vapor. This vapor mixed with blue litmus solution changes the color from blue to red. In place of the sulphur and oxygen, both of which are used up, we have a new substance quite different from either. This new substance into which sulphur and oxygen are changed is called *sulphur dioxide.*

Now this change is like the change when the charcoal was burned. It is another case in which *two substances are changed into one.* Such a chemical change is called *combination.*

**Decomposition or Combination?** — When ammonia and hydrochloric acid are mixed a chemical change occurs; is it a chemical decomposition or a chemical combination?

*Ex. 9.* — I put a cubic centimeter of strong ammonia water into a tube or bottle, rinse it around to wet the sides, and then pour it out. I put as much hydrochloric acid into

another similar vessel and treat it in the same way. By this means I fill the vessels with the colorless vapor of the two substances. I next bring these two vessels mouth to mouth and hold them one above the other (Fig. 7). Notice and describe


Fig. 7.

The change which the colorless vapors undergo.

The two substances used are colorless, but a new one is made which is seen as a white cloudlike mass, which rolls down toward the bottom of the tube, and will roll back again if the tubes be inverted. In fact, the hydrochloric acid and ammonia *combine* to form one thing, — called *ammonium chloride.*

**Substitution.** — All chemical changes are either decompositions or combinations. But in a great many cases the two kinds take place at once. This is true in the action of zinc and hydrochloric acid.


Fig. 8.

*Ex. 10.* — I have a wide-mouth bottle which will hold 200 cc. and a square of glass or of heavy paper with which to cover it.

Into the bottle I put two or three pieces of zinc, and just cover them with hydrochloric acid. I then close the bottle with its cap of glass or paper (Fig. 8).

Describe the action which sets in.

Feel the bottle, and what effect is discovered?

I now bring a match-flame to the mouth of the bottle while I lift the cover.

Describe the result.

A violent effervescence takes place whenever zinc and hydrochloric acid are brought together. The vessel is heated, and a gas escapes which takes fire with explosion. This gas is *hydrogen.* Heat and hydrogen are two products of the chemical action. Are there any others?

*Ex. 11.* — To answer this question I examine the liquid left in the bottle, when the bubbling of gas has come to an end. I must first *filter* the liquid to rid it of black flakes, which are only impurities of the zinc, and afterward *evaporate* the liquid to recover any solid substance it may contain.

1. *Filtration.* — Cut a square of filter paper, *A*, Fig. 9, about three inches on a side, — a little less than twice the length of the sloping side of a funnel, *f*. Fold the square into a triangle by bringing corners *d* and *a* together. Fold again, bringing corners *c* and *b* together, making the triangle *B*.

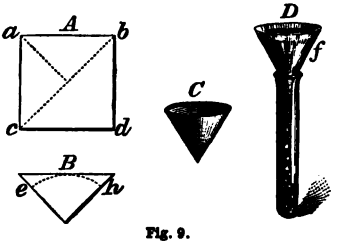
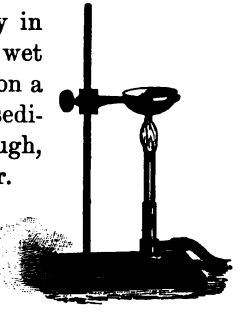


Fig. 9.

Then trim the edges along the circular line *e h*. Open the triangle (leaving three thicknesses of paper on one side and one on the other), and we have a little paper funnel, *C*, which will fit neatly in the glass funnel, *f*. Press it in and wet it with water. Then rest the funnel on a test-tube and pour the liquid with its sediment into it. The liquid will run through, but the sediment will stay on the filter.


Fig. 10.

2. *Evaporation.* — Pour the clear liquid into a porcelain dish and heat it over a low flame, as shown in Fig. 10. It may boil, but should do so very gently. The water will slowly pass away as vapor, but any solid substance which is dissolved in it will stay in the dish.

I will evaporate the liquid left in Ex. 10, and just filtered, down to about one-third its bulk, and then let it

stand until it is cold. A brown solid is now seen, if the evaporation has gone far enough. This brown solid must have been made by the action of the zinc on the hydrochloric acid in Ex. 10. It is called *zinc chloride.*

We began the experiment with zinc and acid; we afterwards find hydrogen, the brown solid, and heat, in place of these two substances.

In all, we have now found three products of the chemical action of zinc and hydrochloric acid. They are heat, hydrogen, and zinc chloride.

Now hydrochloric acid is a compound of the two elements hydrogen and chlorine, and the zinc just takes the place of the hydrogen in it. In fact, the zinc has *decomposed* the hydrochloric acid and *combined* with the chlorine of that substance. The hydrogen was driven out as a gas, which burned quickly while the solid zinc chloride[1] stayed behind in solution.

When one substance takes the place of another in a compound, as the zinc took the place of hydrogen in the acid, the action is called *substitution.* This is one of the most common kinds of chemical action.

**Double Decomposition.** — It often happens that when two things are brought together *both* are decomposed. This is the case with silver nitrate and sodium chloride.

*Ex. 12.* — I take a small crystal of silver nitrate and drop it into a test-tube with 5 cc. of water, and shake it until it is dissolved. In another test-tube I dissolve some common salt, which in chemistry is known as sodium chloride. I now add a drop or two of the solution of salt to the silver nitrate. After seeing,

What happens to show that a chemical action occurs? I go on adding the solution of salt carefully, drop by drop,

[1] The solid is called zinc chloride because it is made up of zinc and chlorine. When pure it is nearly white.

watching to see whether still more of the new substance is made by each addition; and that I may make no mistake about this I shake the tube each time vigorously, and then wait till the solid settles a little before I add the next drops. I do this until a drop of the liquid gives no more of the white solid, and stop with that drop.

If by accident I do get in a drop more than this, I add a drop of silver nitrate solution.

I then filter the mixture as in Ex. 11. The white solid will be left on the filter-paper while a clear liquid will run through.

I next evaporate this clear liquid, as shown in Fig. 10, until but little is left, and then, when it cools, crystals will appear. The new white solid is silver chloride, and the crystals are sodium nitrate.

The silver nitrate and sodium chloride decompose each other, and then the silver and sodium just change places and form the two new compounds named above. The names themselves show this double change, thus: silver nitrate and sodium chloride become sodium nitrate and silver chloride.

In all such cases as this, in which two substances decompose one another and form two new ones, the action is called *double decomposition.* We shall find another example of double decomposition in the action of mercuric chloride and potassium iodide.

*Ex. 13.* — I take enough mercuric chloride in powder to half-way fill the rounded bottom of a test-tube, cover it with 5 cc. of water, and warm it until it is dissolved. I take about twice as much potassium iodide in another tube and dissolve it also in water. I put about 150 cc. of water into a wide-mouth bottle, and add the mercuric chloride solution. I now put in about one-half the solution of potassium iodide, little by little.

Describe the result as it first appears.

And the curious change which soon occurs.

I go on adding the iodide, carefully noticing whether still more of the new substance is made by each addition, and, that I need make no mistake, I wait each time to let the solid settle a little and the liquid clear before I add the next drops. I do this until a drop of the iodide gives no more of the colored solid, and stop with that drop. If by accident I do get in too much, I will add drops of mercuric chloride until the last drop has no effect. The scarlet solid will soon settle and leave a clear liquid above it.

We find that when mercuric chloride and potassium iodide are brought together in solution they yield a substance whose color, at first bright yellow, very soon changes to a fine scarlet. The potassium iodide will give this substance just as long as there is any mercuric chloride in the solution. When a drop fails, we may know that the chloride is all used up. The scarlet *precipitate* (as we call a solid which comes by putting two liquids together) and a solid substance which stays dissolved in the water are the two new things into which the chloride and iodide were changed.

Now the chemists tell us that mercuric chloride contains mercury and chlorine, and that potassium iodide contains potassium and iodine. When we put these two liquids together the chlorine and iodine changed places. Our scarlet solid contains the mercury and iodine, and our white solid contains the potassium and chlorine. In fact, *both* of the old substances were *decomposed*, and the two new ones were made by their parts *combining* in different pairs.

SUGGESTION TO THE STUDENT. — Notice all along how the name of a substance tells us also the names of the simpler substances in it. When we see the name mercuric chloride we may be reminded of the names of its parts, mercury and chlorine; and

in the same way potassium iodide suggests its parts, potassium and iodine. So, too, our yellow solid made of mercury and iodine is named mercuric iodide.

**Definitions.** — A substance which is made up of two or more kinds of matter quite unlike itself is called a *compound.* Mercuric oxide is a compound of mercury and oxygen. For the proof of this see Ex. 5.

The different kinds of matter of which a substance is made are called its *constituents.* Mercury and oxygen are the constituents of mercuric oxide; sulphur and oxygen, of sulphur dioxide (Ex. 8).

By far the larger number of substances are compounds; but there are a few which have never yet been decomposed, and such are called *elements.* Mercury, oxygen, and sulphur are examples of elements. Some seventy of these are now known, and out of this small number all the compounds in nature are made.

But compounds rarely occur *pure.* Substances are mingled together two or more so completely that they seem to be only one, and yet each still has all its properties unchanged. Syrup of sugar is an example. The sugar and the water are simply mixed together without change of properties. Such substances are called *mixtures.* Brine is another example.

All substances are either elements, compounds, or mixtures.

There are two ways of finding out what the constituents of a compound are; one is *analysis,* the other *synthesis.* Any process in which we decompose a substance so as to learn what it is made of is an analysis. Ex. 5 was an analysis of mercuric oxide. Any process in which we make a compound by putting its constituents together is a synthesis. Ex. 8 was a synthesis of sulphur dioxide, and in Ex. 7 we found that carbon dioxide is composed of car-

bon and oxygen by making these two elements combine. This also was an example of synthesis.

**Heat and Chemical Action.** — Heat is very often produced by chemical changes. It was so in Ex. 10. It is so, for another example, in the case of

SULPHURIC ACID AND WATER. — *Ex. 14.* Into a wide-mouth bottle I pour 40 cc. of cold water, and provide a rod of glass, or of wood, with which to stir it. I then pour into it gradually, while I stir it, 40 cc. of strong sulphuric acid.[1] Note the evidence that heat is produced as shown

By handling the bottle.

By inserting a test-tube holding a little alcohol.

These chemical changes need no extra heat to start them, but yield heat as one of the products of the action.

There are other cases in which heat must be used to *start* the change (Ex. 7 and Ex. 8), which once started produces heat enough, and often more than enough, to keep the action going. We touch the wood with a match-flame to *start* a fire, but once begun the burning is kept up by the heat of the chemical action itself.

There are still other cases in which heat must be applied, not only to start the action, but also to keep it going (Ex. 5 and Ex. 6). In these cases heat is not produced, but *absorbed*, by the chemical action.

The fact is that every chemical action is a source of heat or cold; every change in the nature of a substance is accompanied by a change in temperature.

**Electricity and Chemical Action.** — We are told in the study of *Physics*, that if a strip of amalgamated zinc and another of copper are put into a vessel of very dilute acid they will yield a current of electricity.

[1] The heat would be stronger if the water were poured into the acid. *This should never be done.* Whenever strong sulphuric acid and water are to be mixed *always* pour the acid *gradually* into the water while you stir it.

*Ex. 15.* — I first make a little zinc-copper couple in this way: I cut from a sheet of zinc, such as is used under stoves, a strip four inches long by one-half inch in width and bend it squarely at three-fourths inch from one end, *z*, Fig. 11. I then amalgamate it. For this I take a few cubic centimeters of the half-strong acid which was made in Ex. 14, and add to it five times as much water in a saucer. I must also have a little mercury in a test-tube. I first wet the zinc with the acid. I next pour the mercury first upon one side, then upon the other, and rub the whole surface gently with a piece of cloth. The surface of the zinc should now shine like silver; it is amalgamated.

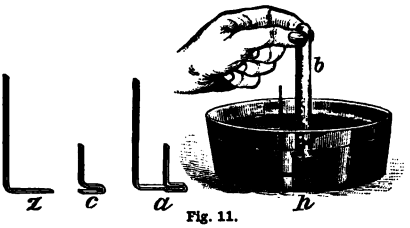


Fig. 11.

I next take a piece of sheet-copper as wide as the zinc and two and a half inches long, and fold one end, as shown at *c*. I press the end of *z* into this fold and make the two fit closely by pressing the fold carefully (not to break the zinc) with pincers. The upright part of *c* is about two inches long and stands facing the upright part of *z*, as shown at *a*, Fig. 11.

To prepare the dilute acid: I measure into the water-pan *p* water enough to be a little deeper than the copper *c* is high, and then add one-twentieth as much strong sulphuric acid.

I shall also need a test-tube: to have it ready I lay one in the water, let it fill and sink.

These preparations all made, I next stand the zinc-copper in the acid water.

A torrent of bubbles rise alongside the copper. What are these bubbles? To catch them I lift the test-tube,

bottom upward, letting no air enter it, and bring its mouth over the top of the copper, as shown at *b*. The bubbles now rise into the tube, driving the water out at the rate of 1 cc. a minute, if everything works well. When the tube is about half filled with gas I light a match, and lift the tube slowly; the water falls out, air takes its place, and as the flame touches the mouth of the tube a sharp report occurs, which says — hydrogen!

The acid water is decomposed and hydrogen is set free. But amalgamated zinc alone will not decompose the acid (try it), nor will copper alone (try it). Is there anything, more than zinc and copper, when they are together in the acid? Yes, there is, as we are told, the current of electricity. The zinc-copper is a source of *electricity*, and the hydrogen is set free by the action of electricity.

It is found that a great many chemical changes can be made by means of electricity. And, on the other hand, a great many chemical changes produce electricity. Like heat, electricity is sometimes an agent and sometimes a product in chemical action.

**Light and Chemical Action.** — Light is also a product of chemical action. The light of our fires and the light of our oil and gas lamps are examples. Remember, also, the burning of sulphur in Ex. 8, and of charcoal in Ex. 7.

Light will also sometimes *produce* chemical changes.

*Ex. 16.* — I first make some silver chloride by filling a test-tube three-fourths full of water, adding a small crystal of silver nitrate, shaking until the crystal is dissolved, and then adding hydrochloric acid drop by drop. The white cloud which rolls down is silver chloride. I now place the tube in strong sunlight.

What change occurs in the color of the chloride?

The change in color slowly from white to dark purple is

a sign that a new substance is being made. In fact, light decomposes the silver chloride. For other examples, and for the explanation of this one, we must wait.

Heat, light, and electricity are *agents* which we can use to bring about chemical actions which in their absence will not occur.

Heat, light, and electricity are also *products* of chemical action. They are just as important to the chemist as are the material products, that is, the new substances, obtained at the same time.

We now go on to examine hydrogen and oxygen gases which we have met with in several experiments.

## HYDROGEN.

Hydrogen may be made by zinc and hydrochloric acid, as it was in Ex. 10. We may make and catch it as follows:

*Ex. 17.* — I make enough dilute hydrochloric acid — one measure of strong acid to two measures of water — to fill the water-pan about two inches deep. I take the graduated cylinder, lay it in the acid, as shown in Fig. 12, and when it is thus filled with the liquid I lift it bottom upward and let it stand mouth down on the bottom of the pan, as in Fig. 13. I next drop a piece of granulated zinc into the pan, and then carry the mouth of the cylinder over it. In all this work I use great care to not let a bubble of air get into the cylinder. The hydrogen bubbles form on the zinc, rise in the cylinder, and push the water down out of it, as shown in Fig. 14, until finally the cylinder is filled with hydrogen.

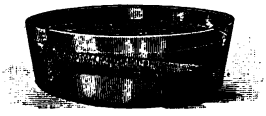


Fig. 12.



Fig. 13.

**To Discover the Properties of Hydrogen.** — I test the gas by lighting a match, then lifting the cylinder out of the liquid rather slowly, mouth still downward, and then quickly bringing the flame to the open mouth. The gas takes fire with a *dull* report. Compare this result with that in Ex. 10.

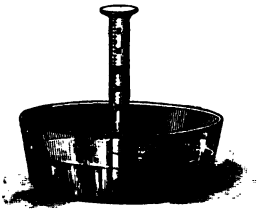


Fig. 14.

*Ex. 18.* — I fill the cylinder again with the gas, and lift it from the water in the same way, then turn it over and let it stand on the table open while I light a match. I bring the match to the mouth of the jar.

Why is there no report?

Which is heavier, hydrogen or air?

What is the color of hydrogen?

*Ex. 19.* — I put 30 cc. of water into the mortar, Fig. 6, and 30 cc. strong hydrochloric acid. I drop into this a piece of zinc, and then place a small funnel, mouth down, over the zinc. I next bring the mouth of a test-tube down over the stem of the funnel, as shown in Fig. 15. The action should be brisk; and if it be so, then, after one minute, I slowly lift the tube — keeping its bottom up — carry it away from the stem, and then bring a lighted match to its mouth. A dull explosion proves that the tube is full of hydrogen.

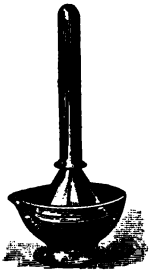


Fig. 15.

The hydrogen rises through the air to the top of the tube, collects there and gradually pushes the air down and out at the bottom. If the explosion is *sharp*, it shows that the tube was only partly filled. Then repeat the experiment, and wait longer.

Other gases which are lighter than air may be collected by this method, called *upward displacement*. Oxygen, which

is heavier than air, is collected (Fig. 5) by *downward displacement.* In Ex. 17 the gas was collected by *displacement* of water instead of air, and this method may be used, whether the gas is lighter or heavier than air.

WHAT CAUSES THE EXPLOSION? *Ex. 20.*—I take a short piece of glass tube, which is drawn out to a small jet at one end, and fix it with a piece of rubber tubing upon the stem of the funnel. I again put dilute hydrochloric acid in the mortar, add the zinc, and place the funnel over it. The gas must come off very briskly; and, if need be to make it do so, I add more *strong* acid. After about half a minute, if the gas is coming fast, and longer if not, so that the air shall be all driven out, I bring a match-flame to the top of the jet on the end of the funnel. The gas takes fire with the usual *dull* explosion, but goes on burning as quietly as a candle. Fig. 16 shows this result.



Fig. 16.

Then is hydrogen itself explosive?

Note also if the flame is pale or brilliant.

Is it a hot flame? Try it with a small wire.

*Ex. 21.*—But what causes the explosion, if hydrogen can burn without it? I will collect a little hydrogen over water, as in Ex. 17. But, instead of *filling* the cylinder with the liquid, I will leave it about two-thirds full of *air*, filling only about one-third its height with the water.

The liquid will soon be driven out, and then the cylinder is filled with a mixture of hydrogen and air. I now lift it from the pan, and at once bring a match-flame below its mouth. A *sharp* explosion follows.

What substances together explode in this case?

Have these same ones been together before whenever the explosion has occurred?

What, then, is the material which explodes?

**What new Substance made by burning Hydrogen?** *Ex. 22.* — To answer this question I must burn pure dry hydrogen and catch the products. I will make the hydrogen as in Ex. 20, but before I light the jet I must dry the gas, and I can do this by passing it over *calcium chloride*, which absorbs water greedily.

I take a "drying-tube," *a*, Fig. 17, put a little cotton-wool in the bulb loosely, drop in small pieces of calcium chloride to fill the tube nearly full, and put a thin layer of cotton over it. I next close the large end with a cork having a hole,[1] through which I crowd the end of the stem of the funnel, *b*, Fig. 17.

I now put several fragments of the zinc into the mortar; afterward place the funnel over the zinc and fix the drying-tube firmly in the clamp of the support *s*, as shown in Fig. 17. This done, I pour dilute hydrochloric acid into the mortar.

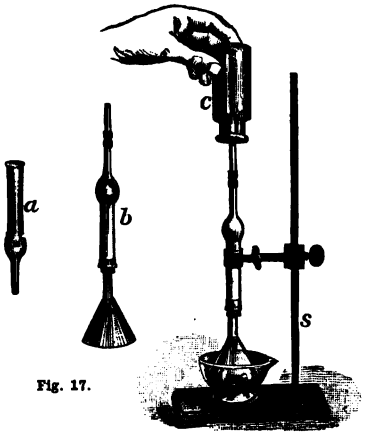


Fig. 17.

Hydrogen is set free. It goes up through the calcium chloride, which takes out the water. The dry hydrogen escapes from the jet above.

After waiting until I am sure that the air is all driven out, I set fire to the dry hydrogen at the top of the tube, and then I hold a glass tumbler, or a wide-mouth bottle, which is clean and dry, over the flame, as shown at *c*.

[1] Holes are made through corks by means of "cork-borers," made for the purpose. They may also be easily made by first running a hot pointed wire through the cork, and then using a round file to enlarge the hole. Each hole should be made to fit its tube very closely.

Notice a deposit of dew on the walls of the bottle; it is nothing but water. This is the substance made by burning hydrogen in air.

What are two other products? See p. 28.

**Description of Hydrogen.** — We have found that hydrogen is a gas without color, and when pure it is also without odor and taste. It is very much lighter than air (Exs. 18 and 19). An equal bulk of air is 14.44 times heavier than this gas. In fact, hydrogen is the lightest of known substances.[1]

Hydrogen unmixed with air burns with a silent flame (Ex. 20), but when mixed with air the mixture burns with explosion (Ex. 21). The chemical action is the same in both cases: water, heat, and light are the products of this action (Ex. 22). There is nothing in this experiment to show what the hydrogen combines with to make the water. Let us remember this.

The flame of burning hydrogen is very hot (Ex. 20). In fact, no other fuel gives so hot a fire as this. One gram of this gas will yield enough heat in burning to boil 344.62 g. of ice-cold water.

A *unit* of heat is so much heat as will raise the temperature of 1 g. pure water 1° C. This unit is called a *calorie*, just as the unit of weight is called a *gram*. The heat of all chemical changes is measured by this unit. The burning of the gram of hydrogen gives 34462 of these units, or 34462 calories.

Hydrogen gas is very seldom found in nature; it occurs sometimes among the vapors which are thrown out of volcanoes. But the compounds of hydrogen are everywhere. Water is only one of them. Nearly all animal and vegetable bodies also contain hydrogen in large quantities.

[1] One liter of hydrogen weighs .0895 g. when its temperature is 0° C., and the pressure of the air 15 lbs. per sq. inch.

QUERY. — A bottle stands mouth downward in water; it is full of gas; how would you decide whether it contains air or oxygen or hydrogen?

## OXYGEN.

Oxygen may be obtained by heating mercuric oxide, as in Ex. 5, or potassium chlorate, as in Ex. 6. The latter way is better, and a mixture of potassium chlorate with "black oxide" of manganese is still better than the chlorate alone.

**To make Oxygen.** *Ex. 23.* — I take 4 g. of the chlorate and grind it to powder in the mortar. To this I add 2 g. black oxide of manganese and mix the two powders. I put this mixture into the side-neck tube, Fig. 4, which I then cork tightly and fix in the clamp, as shown in the cut, Fig. 19.

To catch and hold the gas, I use a set of flasks fitted up as shown in Figs. 18 and 19. A conical flask is provided with a soft rubber stopper with two holes. A long glass tube passes through one of these holes in the stopper, and almost to the bottom of the flask; a short glass tube passes only through the stopper in the other. I use four of these flasks, each holding about 200 cc.[1] and a small bottle, *e*. Into the first one, *a*, I put sand enough to cover the bottom, and then some water, and I also put water into the bottle *e*. I then join them together, as shown in Fig. 19, — the long tube of *a* with the side-neck of the ignition-tube, *i*, and the long tube of each



Fig. 18.

[1] This is a good size for the student's table. Larger ones may be used by the teacher for the class-room, and a flask may be used instead of the tube *i* for larger quantities. A set once fitted up is a sort of general gas-works for the laboratory, for, as we shall see, this method of collecting gas can be used for many other gases beside oxygen.

of the others with the short tube of the one before it, the last rubber tube simply dipping into the water of the bottle.

The gas will be "washed" by bubbling through the water in *a*, and then, too, one can tell by the bubbles in *a* and in *e* whether the gas is coming off fast or slow, and can regulate the heat accordingly. The use of the sand will be seen in Ex. 27.

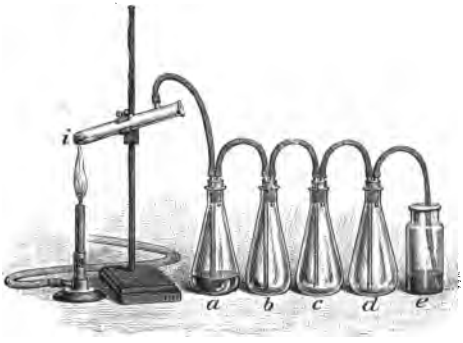


Fig. 19.

All the flasks *must be closed air-tight;* rubber stoppers easily make air-tight joints.

I now go on to make the gas. I make the flame of the Bunsen just high enough to not quite touch the ignition-tube, then lift the lamp and heat the mixture, gently at first, beginning at the part nearest the clamps, and moving the flame to heat the tube uniformly. The decomposition of the chlorate will gradually go on until it reaches the bottom of the tube. If the gas comes off too fast at any time I withdraw the heat for a moment.

The gas will bubble through the water in *a*, and gradually fill that flask, pushing the air in it over into the next one, *b*. When *a* is full, oxygen will go over to the bottom of *b*, and will gradually fill that flask, pushing the air over into *c*. From *c* the air is driven over to *d*, and from *d* into *e*, the oxygen filling them all. One can be sure that the vessels are all filled with this gas by putting a lighted match into the mouth of *e;* the flame will be brightened.

When this test shows that the flasks are full of oxygen,[1] or when all the chlorate is decomposed, if in any case too little is used, I withdraw the flame, and take off the rubber tubes from all the flasks.

**The Chemical Change.** — Potassium chlorate is made of potassium, chlorine, and oxygen. It gives up its oxygen when heated, while the other two elements are left in the tube combined with each other. The manganese oxide is unchanged, and yet its presence is very useful, because it compels the chlorate to give up all its oxygen more steadily and with less heat. But just how it does this is not known. This is one of the cases in which a substance seems to act by its *presence* simply. A chemical action like this, that is, one due to the presence of a substance which remains unchanged, is called *catalysis.*

**To discover the Properties of Oxygen.** — The flasks, used in Ex. 23 are full of oxygen.

What is the color of this gas?

The bottle *e* is also full: what is the odor of oxygen?

The bottle has now been standing some time open: is it still full of gas? If so, then which is heavier, oxygen or air?

*Ex. 24.* — I remove the stopper from flask *d*, hold the flask bottom up for two minutes, then stand it on the table and insert a lighted splinter of wood.

Does the gas remain in the flask?

Then which is heavier, oxygen or air?

Was this also shown by the bottle *e?*

Oxygen and the Match. *Ex. 25.* — I wind a small copper wire (No. 20) around the white end of a burned match. I then heat the black end to a glow, and lower it a little way into the bottle *e*. The glow promptly becomes

[1] The amount of chlorate used (4 g.) will yield more than enough to fill these four flasks.

a flame. I blow the flame out, and again lower the spark into the oxygen; the wood is promptly relighted. This may be repeated many times with the same flask full of gas.

Oxygen and Carbon. *Ex. 26.* — I take a piece of charcoal — of the bark, if I can get it — and twist the end of a piece of fine copper wire around it for a handle. I then heat one corner of the charcoal until it glows and quickly lower it into flask *c*. After the burning is over I pour a little lime-water into the flask and shake it well.

What effect has oxygen on the burning of charcoal?

What effect is produced on lime-water afterwards?

What substance must be present to do this? (See p. 17.)

Oxygen and Iron. *Ex. 27.* — I take a piece of small iron wire (No. 23), and bind a piece of match to one end of it. I then set fire to the match and place it in the mouth of flask *a*, lowering it slowly as it burns away. The burning wood heats the iron until it too takes fire, and it then burns with surprising brightness. A *slender* watch spring would burn with still greater beauty. The sand catches the melted globules of iron as they fall, and may save the glass from being broken.

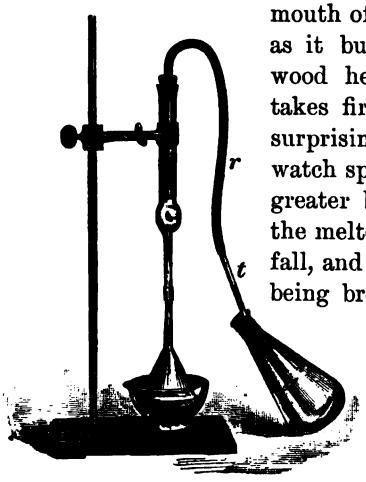


Fig. 20.

Oxygen and Hydrogen. *Ex. 28.* — Finally in the flask *b* I will burn a jet of hydrogen. I make the dry hydrogen as in Ex. 22. The only change in the apparatus is the addition of the tube *r t*, Fig. 20; *r* is rubber and *t* is glass, the lower end of which I held in the flame until the

hole was nearly closed. Out of this small hole the gas comes in a small jet. *After the air has been all driven out* I set fire to the hydrogen and then plunge the small flame down into the oxygen in the flask, as shown in the cut.

What change in the appearance of the flame?

What is the new substance on the walls of the flask?

What are the two constituents of this substance?

What must have combined with the hydrogen in Ex. 22?

Where did the hydrogen there get it?

**Description of Oxygen.** — We have found that oxygen has neither color nor odor nor taste. It is about 1.1 times heavier than the same bulk of air. Oxygen combines readily with many things. It does so very rapidly when heated, and the chemical action yields both heat and light as well as new substances. A chemical action which yields both heat and light is called *combustion.* The heat and light of all common fires are due to the action of oxygen, which abounds in the air.

Bodies which burn at all in air will burn with much more vigor in oxygen alone, as did the match and the charcoal in Exs. 25, 26. And many things which do not burn in air will burn freely in this gas, as did the iron in Ex. 27.

When a thing is combined with oxygen it is said to be *oxidized,* and the new substance is called an *oxide.* Iron is oxidized when it burns in dry oxygen, and the new substance made is the iron oxide. Carbon burned in oxygen (Ex. 26) is oxidized and carbon dioxide is produced, which will show its presence by whitening lime-water.

But most substances must be heated before they will oxidize rapidly: neither wood nor coal nor iron will burn unless first made much hotter than they can ever become by exposure to the greatest summer heat. And yet the oxygen of the air is all the time acting upon many things. Wood decays and iron rusts; these effects are due to

oxygen. But they are not produced quickly. Substances are oxidized, at ordinary temperatures, slowly.

**Oxygen in Nature.** — About one fifth part of all the atmosphere is oxygen gas, and eight-ninths the weight of all the water of the earth consists of this element. In the bodies of animals and of plants oxygen is found in large proportions, and in the rocks immense quantities are combined.

QUERY. — A bottle is full of a colorless gas; how would you decide whether it is air or hydrogen or oxygen?

**Ozone.** — Oxygen is very strangely changed by the action of electricity. If electric sparks are sent through oxygen, the gas will be found to have a very strong and peculiar smell, a little like that of burning sulphur. This strong-smelling oxygen will tarnish silver, which remains bright in oxygen that has not been electrified, and it will do much other chemical work which common oxygen cannot. Two cubic centimeters of it weighs as much as three cubic centimeters of oxygen. In weight and odor and in chemical activity this electrified oxygen is as different from common oxygen as if it were another substance.

And yet it can be nothing but oxygen. For the electric sparks cannot add anything to the pure oxygen they pass through, and so in going through it they cannot make a compound of the oxygen. And if the oxygen was decomposed by the sparks, there would be at least another constituent set free beside this one; but there is not.

We must confess that oxygen has two sets of properties; the gas may exist in two distinct forms. The heavy, active, strong-smelling oxygen is called *ozone.*

Just how the electricity changes oxygen to ozone is not known. But one thing is settled by experiment, and that is that the oxygen is condensed; three volumes of oxygen

will make only two volumes of ozone. The electricity condenses the oxygen in making ozone.

Ozone is also made when acid-water is decomposed by electricity (Ex. 15). It is found in small quantities in the atmosphere, especially after thunder-showers, on account of the electric discharges, and its presence tends to purify the air; ozone oxidizes the impurities quickly and makes them harmless.

Oxygen is not the only element[1] which can exist in two forms, as we shall see. This property of an element is called *allotropism*. Ozone is said to be an allotropic form of oxygen.

### *EXERCISES IN CHEMICAL ACTION.*

**1.** Study the action of hydrochloric acid on sodium carbonate by experiment. Proceed as follows:

*1.* Bring the two substances together in a test-tube, and describe the action which takes place.

*2.* Examine the gas which is given off to find out what it is.

Note its color and its odor.

Test it with a flame; is it oxygen or hydrogen?

Is it heavier or lighter than air?

Collect a little of the gas in another test-tube. To do this, put a little of the carbonate into a side-neck tube, and put the rubber tube from this down into a test-tube. Make a very dilute acid by putting a cubic centimeter into the graduated cylinder and filling with water up to 25 cc.; this dilute acid will not work quite so inconveniently fast as the strong. Now pour three or four cubic centimeters upon the carbonate, and cork the side-neck tube quickly. When the action is nearly ended, add the acid again. Next pour a little lime-water into the test-tube with the gas and shake it well.

[1] See definition of element on p. 24.

What gas does it prove to be?

*3.* Examine the liquid which remains in the side-neck tube to learn whether it contains any solid in solution. Filter it (Fig. 9), if it is not already perfectly clear. Then evaporate it (Fig. 10) until it is dry.

Compare the solid found with the carbonate used.
You can recognize this new substance by its taste.

*4.* Write out a short statement of the facts which you have discovered about the action of hydrochloric acid on sodium carbonate.

**2.** Study the action of hydrochloric acid on "Baking soda." Do this by experiments in just the same way as in Exercise 1.

What are the two products of the action?
*What, then, is "Baking soda"?*

**3.** Study the action of dilute sulphuric acid on iron. Do this by bringing the acid in contact with some very small nails in a test-tube, closing the mouth of the tube, but not quite air-tight, with the finger. Then

Find out whether the action is hastened by heat.
If a gas is produced, *test it,* and name it.
Find out by evaporation whether a solid is made also.

And finally write out a short statement of all the facts which you have discovered about the action of dilute sulphuric acid on iron.

**4.** What is the difference between a physical change and a chemical change? Between experiment and observation? Between analysis and synthesis? What are elements? What are compounds? What are mixtures? What, besides new *substances,* may be produced by chemical action? What is combustion? When is a substance said to be oxidized? What is an oxide?

## THE CHEMISTRY OF COMBUSTION.

In ancient times only four elements were supposed to exist, and these were fire, air, earth, and water. It is now known that neither one of these is an element; and as to *fire*, we know that it is not even a substance at all. Fire is the light and heat produced by chemical action. The chemical action which produces fire is called *combustion*.

Between what elements is this chemical action taking place, and what new substances are made by it? These are the questions we now set out to answer.

**Combustion of a Candle.** *Ex. 29.* — I bring a clean and dry bottle down over the flame of a burning candle and hold it there for a little while, as shown in Fig. 21.

What is the effect on the flame?

What gathers on the walls of the bottle?

Will lime-water be whitened if shaken in the bottle?

Thus we find water and carbon dioxide both made by the burning candle. But we know that it takes hydrogen and oxygen to form water (Ex. 28), and also that the air can furnish the oxygen (Ex. 22). So, when the candle burns, the air must furnish oxygen, and the candle must furnish hydrogen to form the water which it produces.



Fig. 21.

In Ex. 26 we found that carbon dioxide is made of carbon and oxygen. The carbon dioxide of the candle-flame must contain these same elements. The oxygen could come from the air; the carbon must have come from the candle. Clearly the candle must contain carbon and hydrogen among its elements. And when it burns,

the chemical action is between oxygen of the air and these elements of the candle.

**Combustion in other Cases.** *Ex. 30.* — Burn a splinter of wood, a small roll of paper, a bit of cotton on a wire and moistened with alcohol, and a small jet of gas from the Bunsen lamp, in dry bottles, and see whether you find the same two products in each case.

In the burning of a candle, of wood, of paper, of alcohol, of gas, — and in fact of all other common fuels, — the combustion is simply the combination of oxygen with the constituents of these substances.

Carbon and hydrogen make up by far the larger part of all the substances which are used for fuel, and when the fuel burns carbon dioxide and water are always the chief products. When wood burns it is first decomposed by heat. Its carbon and hydrogen then take oxygen from the air and make carbon dioxide and water-vapor, and these two gases pass away in the smoke.

Why did the flames all go out in the bottle?

The black part of smoke is the carbon which goes off without burning.

A smoky flame is one that gets too little air. A lamp without a chimney smokes, but with a chimney the flame is clear and bright. It is so because the hot chimney makes a current of air sweep past the flame all the time, and this large quantity of air gives oxygen enough to burn the carbon completely. In every common fire much fuel is wasted as smoke, because the furnace is not built in a way to furnish air enough to burn up all the carbon.

Combustion is a mutual chemical action, generally between oxygen and some other substance, and which, when rapid enough, evolves heat and light.

When a substance will burn in air, it is said to be *com-*

*bustible:* the air at the same time is said to be the *supporter of combustion.* But really there is no difference in the part played by the two things in the action: the chemical change is mutual.

**Heat a Product of Combustion.** — Fuel is burned for the sake of the heat it can give, and not for the sake of the new compounds which it yields. The hottest kind of flame is that of hydrogen burning with oxygen. It is called the *oxyhydrogen flame.*

Fig. 22 shows how this flame is obtained. The two gases are in separate bags, *H* and *O*, or sometimes in iron cylin-

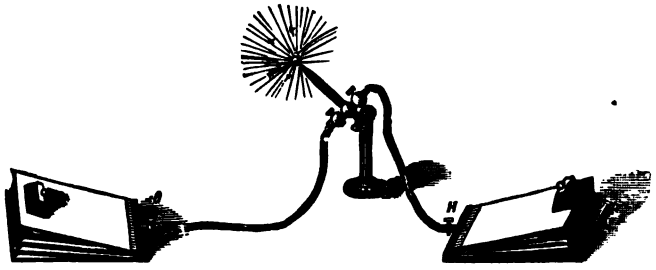



Fig. 22.

ders. They are pressed out of these through separate tubes into the "oxyhydrogen jet," where they mix just before they reach the fire at the end of the jet.

The two gases unite to form water, which goes away as vapor into the air. This chemical action is the source of a heat so intense that wires or strips of iron, steel, copper, zinc, and other things that do not burn at all in common fires, will burn in it almost as fast as a cotton thread will burn in a lamp-flame.

If we burn hydrogen in air instead of in oxygen the heat is less intense because the chemical action is hindered by the large quantity of nitrogen which air contains.

The *quantity* of heat will depend on the quantity of hydrogen which burns. Two grams of hydrogen will yield exactly twice as much heat as one gram, no matter whether it burns slowly or swiftly. But if it burn swiftly, *more heat* will be given *in the same time,* and the heat will be more *intense.*

If a gram of hydrogen burns in the air instead of in oxygen it will still give the same quantity of heat, and yet the heat will not be as intense even if it burn in the same time, because some part will be used in heating up the large amount of nitrogen in air.

We may say the same things of the heat when carbon or any other element is burned. The quantity of heat will always be the same for the same weight of the element burned, while the intensity of the heat will depend on the time it takes to burn it.

But there is this difference, the gram of hydrogen will give more heat than the gram of carbon. There is just a certain amount of heat which the burning of a gram of each element will give, but this quantity is not the same for any two of them.

**Heat Required to Start Combustion.** — A jet of gas, such as that from a chandelier, for example, escaping into the air, shows no signs of "taking fire," but touch it with a match-flame, and it instantly springs into vivid combustion. What has the match-flame done? It has simply *heated the gas.* Illuminating-gas will not burn until it has a temperature of about 1000° F., and when the fire of the match has heated the jet up to this temperature it bursts into flame.

The temperature at which a substance begins to burn in air is called its *kindling-point.* The kindling-point of most of our ordinary fuels is about 1000° F., but some other things begin to burn at a much lower temperature. Phos-

phorus, for example, kindles at a temperature little higher than that of our fingers when we handle it.

The kindling-point of a substance is the temperature at which it will begin to burn.

In lighting a gas-jet the match-flame is needed only to heat the gas up to its kindling-point.

**All Flames are Gas Flames.** — Let the wick of an alcohol-lamp be uncovered; no signs of flame are to be seen, but touch it with a lighted match, and very quickly an alcohol flame appears. Now, the wick, to begin with, is wet with liquid alcohol. Then the heat of the match *changes this liquid into vapor*, and afterward, quickly *heats this vapor up to its kindling-point.* When this double work is done the flame appears.

The alcohol is in a gaseous form when it burns with a flame.

We will suppose, next, that we have a candle which has been lighted and partly burned on some previous occasion. Its wick is saturated with cold and solid wax. We touch it with a match-flame. We notice that it takes more time to fire it than it does a spirit-lamp or a gas-jet. The match has more work to do. It first *melts the wax;* it next *changes it into vapor;* and then, finally, *it heats the vapor up to its kindling-point.* Not until this threefold work is done does the candle-flame appear. The wax is in the form of gas when it burns with flame.

The wax of the candle, and the alcohol of the lamp, are changed into gases before any flame is seen, and this is true of other fuels also. *Whatever burns with a flame must be at that moment in a gaseous state.*

Wood burns with flame, because it is first decomposed by the heat. Gases are formed, and the burning of these gases, and not of the solid wood, produces the flame.

Hard-coal is made up almost entirely of *solid* carbon,

which no furnace-heat can change into *gas*. As there are no gases first made by the heat, so there can be no flame produced in the burning. Hard-coal burns with a steady glow without flame.

This is true if there is plenty of air for all the carbon, but sometimes there is not, and then carbon dioxide is formed, at first, as usual. But this afterward shares its oxygen with more carbon and becomes *carbon monoxide*. This carbon monoxide is a gas, and burns with a blue flame, which may be often seen playing over the surface of a hard-coal fire.

**Light a Product of Combustion.** — The very hot flame of hydrogen gives very little light, but if I hold a small iron wire in this flame the wire will quickly glow with light of a bright-red color. If a piece of lime is held in the oxyhydrogen flame it will shine with a dazzling brightness. The light made in this way is the well-known "lime light," also called the oxyhydrogen light.

In both these cases the light is made by heating a solid substance which will not melt nor become a gas. Any flame which contains a solid substance which will not melt, nor become a vapor when heated, is a light-giving flame. Is there such a solid substance in a candle or a gas flame?

*Ex. 31.* — I take a square of clean dry window-glass and hold it for a very short time across a candle-flame, just below the tip of it. I must remove it before it gets hot.

What products are deposited on the glass?

*Ex. 32.* — I close the holes in the tube of the Bunsen burner and the flame at once becomes bright. I next press a clean dry cold glass down upon it as I did upon the candle-flame before, and notice

Whether the same products are left on it.

*Ex. 33.* — Open the holes of the Bunsen burner and let the air enter. It *mixes* with the gas in the tube, and there is then oxygen enough among the particles of gas to satisfy both the hydrogen and the carbon at once.

Compare the light with that when the holes are shut.

The candle-flame and the gas-flame both leave a black coat upon the cold surface of the glass, and outside this spot a ring of dew may be seen. The black substance is carbon. The wax of the candle contains hydrogen and carbon. Now, when these two are offered to oxygen the oxygen will take hydrogen first. This is one fact. Another is, that *carbon is a solid* which will not melt.

We can now see how the light of a common flame is made. The burning substance is decomposed into hydrogen and carbon. The oxygen of the air combines with the hydrogen first, and produces water, and great heat. This heats the particles of the carbon *white-hot*, so that they shine with a bright light.

But the next moment this white-hot carbon unites with oxygen of the air, and is changed at once into invisible carbon dioxide.

If the hydrogen and oxygen both burn at the same instant little light is given by the flame. The carbon changes to carbon dioxide as fast as the hydrogen does; its particles do not remain free long enough to shine. This is also the reason that the Bunsen flame is smokeless. The oxygen of air is mixed all through the gas and burns the carbon as fast as it is set free.

But the heating of solid particles in a flame is not the only cause of the light. It has been found that some of the light comes from dense gases in the flame as well as from the solid particles. This explains why a lamp-flame is not so bright on the top of a high mountain as it is at the base. The gases in the flames are denser at the base, where

the atmosphere is heavier, than they are at the top of the mountain.

**A Common Flame is Hollow.** *Ex. 34.* — I lay the stick of a common match right across a good candle-flame, just above the top of the wick, and leave it there only long enough for the flame to scorch it. Where it is not scorched of course there is no fire.

Where is the fire, as shown by the stick?

*Ex. 35.* — I take a square of paper and press it down upon the candle-flame almost to the top of the wick and take it away again just as soon as I begin to see the upper surface blacken, as shown in Fig. 23.

Where is the fire, as shown by the charred paper?

*Ex. 36.* — The larger flame of an alcohol-lamp is better for the last two experiments than the candle. I plunge the head of a match into the dark center above the wick. The wood of the match burns in the edge of the flame, but the head of the match in the center does not.

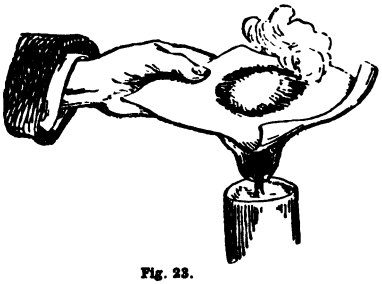


Fig. 23.

What is lacking, heat or air, or both?

In a candle or lamp flame the combustion goes on only around the outside, — in other words, the flame is hollow. There is no oxygen in the center of the candle or the gas flame. The dark space there is filled with the hot vapor of the wax or with gas, and the combustion goes on only where the air is in contact with the outside of this vapor.

The burner of a chandelier is made so as to spread the illuminating-gas out into a fan-shaped sheet. This brings

a larger surface to the air and makes more light, but the chemical action is only on the outside of this thin sheet. The argand oil-burner does the same thing in another way. Its wick is thin and cylindrical, and air is made to pass up through the inside of it. The inside and outside together form a large surface. And then by using a chimney a draught is made by which more air than otherwise must pass up over the surface of the flame. The best light is produced by securing a full supply of air and a large surface to the flame.

QUERIES. — What will be the effect of cooling a flame to a temperature below the kindling-point?

If a large piece of flat iron or stone is laid across a flame, as the paper was laid in Ex. 35, the flame will not touch it; a thin space between them can be seen (try it). Why does the flame not touch the solid?

Does a flame actually touch the bottom of a kettle in which water is being heated? Why?

Why will a flame not pass through wire gauze? But what if the gauze becomes red-hot?

Why does blowing a candle quench the flame? Why does blowing a fire make it burn more briskly?

## THE CHEMISTRY OF WATER.

**Analysis and Synthesis.** — We have seen that there are two ways of finding the composition of a compound: one is called *analysis*, the other *synthesis*. In analysis we decompose the compound in such a way as to show what it is composed of, while in synthesis we combine the constituents in such a way as to show what they make.

Experiment 28 was a synthesis of water, because we

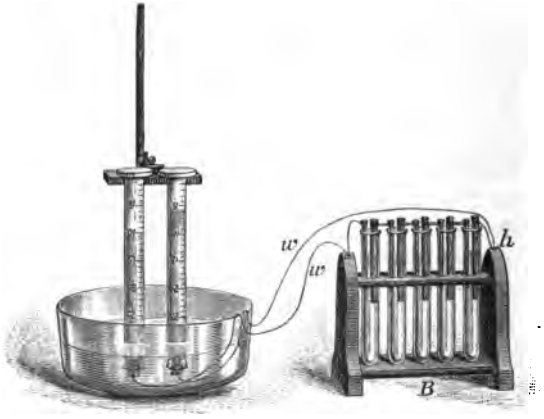



Fig. 24.

brought hydrogen and oxygen together and found that they produced water when they combined. That experiment proved that water is composed of hydrogen and oxygen. But this is not enough. We wish to know whether it makes any difference if we use more or less of these elements, and if it does, then we wish to know just *how much* of each is needed. Now that we know, by synthesis, what the elements of water are, we will try, by analysis, to find

out whether there is any particular quantity of each, and if so, then how much the water contains.

**Analysis of Water.** *Ex. 37.* — How can we decompose water? We have already found in Ex. 15 that electricity will set hydrogen free from acidulated water, and we will try electricity for our purpose now.

APPARATUS NEEDED. — We must have a battery, *B*, to furnish the electricity; two wires, *w w*, to carry the electricity into the water; two graduated cylinders, to catch the gases, the water-pan and the support. All these are shown in Fig. 24.

THE BATTERY. — Two cells of any good battery will decompose water slowly, and a larger number more rapidly. If a battery is not at hand, one can be easily made as follows:[1]

TO MAKE THE TEST-TUBE BATTERY. — Take 20 inches of round "carbon pencil," used in small electric lamps. It should be about $\frac{3}{16}$-inch diameter, and cost a few cents. Saw this into five lengths of four inches each.

Find five nails, each about the same length as the carbons, and place them in a dish of dilute sulphuric acid. The acid will dissolve off the smooth, hard surface of new nails, so that they will act quickly when put into the battery-fluid by and by.

Get some flexible copper wire, — No. 18 is a good size to wind easily, — and finally also get the rack of test-tubes and a pair of pincers.

Take a piece of the wire about four inches long, and wind one end of it around the end of a carbon rod twice, as tightly as it can be drawn, and then, lapping the short end over the wire, twist them with the pincers. This makes a close, firm joint. See Fig. 25, *c*. Then take one

[1] My test-tube battery is cheaply and easily made, and works vigorously. It is fairly constant for an hour and a half. It will yield hydrogen from acid-water at the rate of 1 cc. per minute.

of the nails, *n*, and wind the other end of the wire around it just below the head, drawing the wire as tightly as possible. Roll the wire upon the nail until the carbon and nail are just far enough apart to let the couple hang close against the inside of two tubes when they stand in the rack, as shown at *B*, in Fig. 25 and Fig. 24. Make four of these couples and hang them in the tubes, so that there will be a carbon and a nail in each, except the first and the last. The carbons and nails *must not touch one another*. Then fix one carbon without a nail, and one nail without a carbon, *C*, Fig. 25. Put the carbon in the tube with the lone nail, and the nail in that with the lone carbon, and put the ends of their wires into the small holes *h*, Fig. 24, made with an awl in the top of the rack.

THE WIRES. — Cut two strips of platinum foil, each an inch long and a little less than half an inch wide, and two covered copper wires, say twenty inches long. Make one end of each wire very bright; crowd it through a *small* hole near the end of a platinum strip, bend it back on the other side, and press the loop carefully but tightly together to hold the platinum firmly. Now lay a thin piece of solder on the wire where it touches the platinum, moisten it with a drop of hydrochloric acid, and then hold it in the blue flame of the Bunsen lamp until the solder melts, and no longer, lest the solder corrode and ruin the foil. In this way the wire and foil will be "soldered" together.

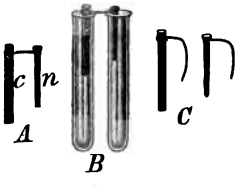


Fig. 25.

The *junctions* of the platinum and wires, and so much of the *wires* themselves as will be in the water, must be *well covered* with paraffine. Melt some paraffine over a *gentle* heat, and before the liquid has become very hot put the lower end of the platinum and wire into it.

See the foot-note on page 54.

THE GRADUATED CYLINDERS. — Use two of the graduated cylinders shown at *a*, in Fig. 1. To support them: saw two slots in a piece of thin board (Fig. 26, *b*), and then fasten it in the clamp of the support *f*, Fig. 5. The cylinders will hang bottom upward through these slots, as shown in Fig. 24.

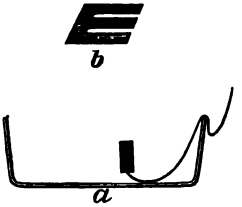


Fig 26.

THE WATER. — Pure water will not let electricity go through it, but if it contains some sulphuric acid the electricity will go freely. Use about one-thirtieth as much strong acid as water, and fill the water-pan deep enough to cover the tops of the platinum strips half an inch when they stand upright, as shown in Fig. 26, *a*.

THE BATTERY FLUID. — Dissolve 35 g. of powdered potassium dichromate in 200 cc. of hot water. Then add 24 cc. of strong sulphuric acid *very slowly*,[1] all the time stirring the liquid. Use it when cold.

THE EXPERIMENT. — Fill the cylinders by laying them in the pan of water, and then, lifting them bottom up, carefully rest them in their support, as shown in Fig. 24.

Not a bubble of air should remain in either one.

Bend the wires to bring the platinum strips up under the mouths of the cylinders, and fix them in place by bending them tightly over the edge of the pan. See Fig. 26, *a*, and Fig. 24.[2]

[1] Great heat is produced by adding the strong acid, and there is danger that drops of the hot liquid will fly out of the vessel. Let the acid run down the side of the vessel in a very small stream while the *red liquid is kept in motion.*

[2] Or, better, the wires may be passed up behind the cylinders through small holes in the wooden support. The wires may be wedged tightly in these holes, and the platinum strips will then be held in place firmly while the wires are being handled during the experiment.

Next, fill the test-tubes nearly full of the battery fluid, put globules of mercury into the holes *h*, and finally put the ends of the wires from the water-pan into these holes, noting by the watch the time when the last is inserted.

Bubbles of gas instantly form on the platinums, break away, and rise into the cylinders; see that none escape outside.

Note the number of cubic centimeters of gas in each cylinder:

At the end of 2 minutes.

At the end of 5 minutes.

At the end of 10 minutes.

Do you find the larger quantity over the platinum which is joined to the carbon or to the iron of the battery?

**Which is Hydrogen and which is Oxygen?** *Ex. 38.*— I test the gases with a match-flame. Holding a lighted match in my right hand, I grasp the cylinder which is over the plantinum of the wire which comes from the iron of the battery, and, closing its mouth as well as I can with my thumb, I lift it out of the water, turn it quickly mouth upward, and at the same time bring the flame to its mouth. The gas burns with explosion.

I next test the gas in the other cylinder in the same way; the match burns with unusual brilliancy, or a long splinter of wood with a spark of fire thrust down into the gas bursts into flame.

What are the relative volumes of hydrogen and oxygen?

**The Facts.** —The current of electricity decomposes the acid water. Hydrogen and oxygen are the only substances produced. And there are always twice as many cubic centimeters of hydrogen as of oxygen.[1]

[1] Oxygen combines with copper, and if the wires and solder are not *completely covered* with paraffine where they touch the liquid, some of the oxygen will be used up in this way.

These facts mean that water is made up of two measures of hydrogen and one measure of oxygen.

**A Source of Doubt.** — But the water was not pure, and some part of these gases may have come from the acid in it, or have been used up by it. So this analysis alone does not *prove* the composition of water beyond doubt. In fact, it is only one source of evidence, while there are many others.[1] We may mention one. By *measuring* the hydrogen and oxygen gases, and then passing an electric spark through the mixture, they are made to combine, and in every case it is found that just twice as many cubic centimeters of hydrogen as of oxygen have been used to produce water.

*All the facts* when taken together *prove* that pure water is composed of

Two volumes of hydrogen and one volume of oxygen.

**Composition by Weight.** — But hydrogen is much lighter than oxygen. In fact, it is found that one measure of oxygen weighs eight times as much as the two measures of hydrogen. So that *by weight*, pure water is composed of

One part of hydrogen and eight parts of oxygen.

One ninth part of any weight of water is hydrogen, and the other eight ninths of it is oxygen.

Then how much of each must there be in 100 grams? One ninth of a hundred, or 11.11 of hydrogen, and eight ninths of a hundred, or 88.89 of oxygen. The chemist writes this composition of water thus:

| | |
|---|---|
| Hydrogen . . . . . | 11.11 |
| Oxygen . . . . . | 88.89 |
| Water . . . . . | 100.00 |

And this is called the *percentage composition* of water.

**Composition by Volume.** — We have just seen that water when it is decomposed yields just twice as many

[1] See Roscoe and Schorlemmer, pp. 204, 212.

cubic centimeters of hydrogen as of oxygen. When we had 20 cc. of hydrogen we had also just 10 cc. of oxygen. But we did not notice *how much water* gave us the 20 cc. of one and the 10 cc. of the other. The fact is that it takes only about $\frac{1}{35}$ part of 1 cc. of the liquid, and this is so small a quantity that we do not see the loss of it in the experiment.

But while 20 cc. of hydrogen and 10 cc. of oxygen will make so little of the *liquid* water, they will of course make a great deal larger volume of water *vapor*. It has been found that they will make just 20 cc. The volume of the water, *in vapor*, is just the same as the volume of the hydrogen alone. The fact is that

Two measures of hydrogen and one measure of oxygen make two measures of water-vapor. Or *three* volumes of the constituents are condensed to *two* volumes of the compound.

Make a note of this curious fact.

**The Constant Composition of Water.** — Water has been analyzed over and over again with the same result. The synthesis of water also has been repeated a great many times, and the same proportions of the same elements have always been found. It is, therefore, quite certain that pure water is always made up of the same elements and in the same proportions by volume and by weight.

**The Constant Composition of other Compounds.** — We may mention hydrochloric acid. It has been proved by both analysis and synthesis that this acid is always made up of hydrogen and chlorine, in the proportions of 1 part by weight of hydrogen to 35.5 parts of chlorine. Its composition is as constant as that of water.

And common salt, from whatever source it comes, always contains the same two elements, — chlorine and sodium, and always in the proportions of 35.5 parts of chlorine to 23 of sodium.

**The Law of Constant Proportions.** — So many compounds have, like these just named, been found to have a constant composition, that the chemist is quite sure that all compounds are alike in this respect, and he states this conclusion as follows:

*Any compound is always made of the same constituents and in the same invariable proportions.*

This important statement of fact is known as the law of constant proportions.

**Water in Nature.** — Hydrogen and oxygen are the only constituents of all pure water. But pure water is not to be found in nature. Even the raindrop, which has never touched anything but the air through which it falls, is not pure, and it is still more impure after it has touched the earth. Why is water always impure? And what are its impurities? These are the questions which we will try next to answer.



Fig. 27.

*Ex. 39.* — I fill a bottle three-fourths full of clear water. I cover it with a piece of muslin loosely, and bind the cover in place by a string around the neck. I put half a teaspoonful of powdered cochineal on the cover, and then pour some clear water slowly upon it. The water very soon trickles through the cover and falls into the water below. But instead of being clear and colorless, it falls from the cover in a *crimson* stream. Why?

*Ex. 40.* — I repeat Ex. 39, but in place of the cochineal I use some powdered copper sulphate. The stream which falls into the water is *blue*. Why?

*Ex. 41.* — I prepare a flask of hydrochloric acid gas in this way: I put 40 or 50 cc. of strong hydrochloric acid into a side-neck flask, *f*, Fig. 28, and fix the flask in the clamp of the support. I join the side-neck to the long tube of the

flask *a*, and the short tube of *a* with the long tube of *b*. And now, all the joints being tight, I gently heat the acid.

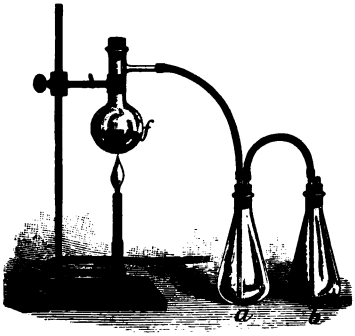


Fig. 28.

Hydrochloric acid gas will go over, driving the air before it, until the flasks are full. When the gas issues freely from *b* it will make itself known; then I withdraw the flame, and at once take both rubber tubes from *a*. I then take the stopper from the flask, and at the same time cover its mouth with the palm of my hand. I now turn the flask bottom upward and lower its mouth into the water of the pan, as shown in Fig. 29. The water rises — it should rise quickly — into the flask. Why?

Cochineal readily mixes with water, and gives it a crimson color. And yet not a particle of cochineal can be seen in the crimson liquid. It is divided into pieces too small to be seen, and these pieces are uniformly scattered, giving color to every part. The cochineal was *dissolved* by the water. The red liquid (Ex. 39) is a *solution* of cochineal in water.


Fig. 29.

Water also dissolves copper sulphate (Ex. 40), and the solution is blue.

Salt is also *soluble* in water; but its solution, brine, is without color. A vast number of other solids are more or less soluble in water. Some of them give it color; others give no visible sign of their presence. The most colorless water may hold many and a great deal of these soluble bodies. They are its impurities. And all water that has

been in contact with soil and rock holds more or less of these impurities.

Gases also are soluble in water. Hydrochloric acid gas dissolves in large quantity (Ex. 41); 100 cc. of water will take up 45,000 cc. of this gas at a temperature of 15° C., — more if it be colder and less if it be warmer. Hydrogen is slightly soluble in water, oxygen a little more so; 100 cc. of water will take only about 3 cc. of oxygen. Water absorbs the gases of the atmosphere. These also are impurities in all natural water.

**Mineral Waters.** — When water holds enough of any one thing in solution to give it a peculiar taste it is called *mineral water.* Such a water receives the name of the substance in it, and so does the spring from which the water issues. Sometimes a spring yields water in which compounds of iron are dissolved; it is then called an iron spring, or a *chalybeate* spring. Sometimes the compounds of sulphur are present in large quantity; the water is then a sulphur-water.

When water holds much lime or magnesia compounds in solution it is called *hard water;* when nearly free from these it is called *soft water.*

**Drinking-Water.** — Good water for household use always contains air in solution. On standing quietly in a warm place a vessel of water will show bubbles of air clinging to the inside surface. Try it. This air if pure is not only wholesome but helps to make the water palatable. But water will absorb bad gases as freely as the good; drinking-water should not be allowed to stand in bad air.

Water which holds a small quantity of mineral matter in solution is also wholesome, but never when it contains animal substances. Typhoid fever, diphtheria, and some other diseases, are frequent where the water used in households is charged with even very small quantities of animal substances.

Vegetable matter is less dangerous than animal matter, but when much is present it likewise makes the water unfit for household purposes.

Water which holds fine particles, whose impurities are in the form of sediment, may be purified by filtration. On a large scale, the water for cities is filtered through beds of sand and gravel.

But the most deadly impurities of water are often in solution, and these no such filter can take out. Charcoal-filters may remove small portions, but they cannot be trusted to purify bad waters for household use.

**Distillation.** — Pure water can be obtained by boiling common water and catching the steam in a cold vessel. This process is called *distillation.*

*Ex. 42.* — I place 50 cc. of water in my side-neck flask and close it with a cork through which I have pushed the stem of a thermometer, as shown in Fig. 30. To the side-neck I join a long glass tube by slipping a short piece of rubber tube over the end of each. I put the end of the glass tube down into a test-tube, which I lay in the water-pan nearly filled with *cold* water, and finally I heat the flask with a Bunsen lamp.

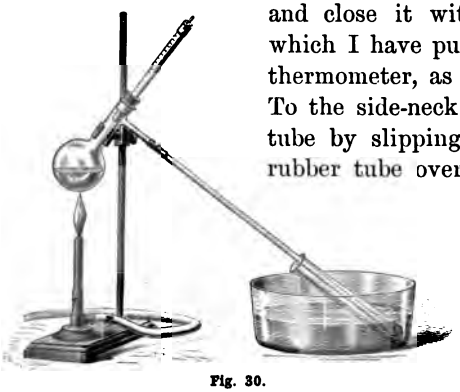


Fig. 30.

Note *every* effect you can see while the water is being heated.

Keep watch of the mercury in the thermometer before the water boils.

Keep watch of the mercury in the thermometer after the water boils. Note the temperature.

What happens in the test-tube?

**The Facts.** — Long before water boils bubbles may be seen escaping from it; these are bubbles of air which was in solution. Afterward the water boils and the steam goes over into the test-tube, where it is at once changed back into water by cold. But it is only the water which goes over as steam; the solid impurities stay in solution in the flask because they need so much more heat than water does

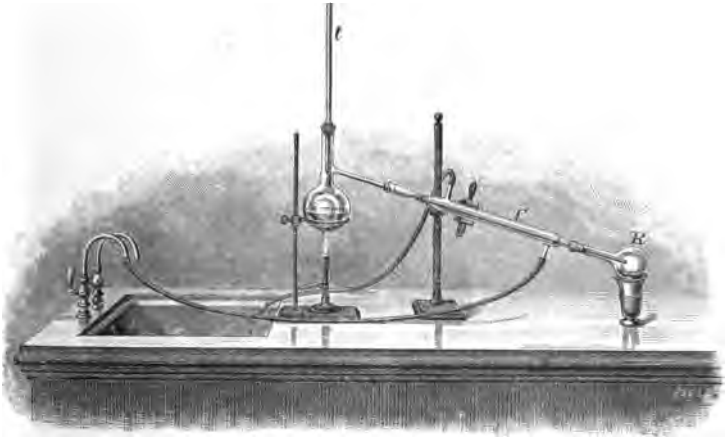


Fig. 31.

to change them into vapor. The water in the tube is nearly pure; it is called "distilled water." And this way to purify a liquid is called *distillation.*

The thermometer shows that in this process the water becomes hotter and hotter until it boils, but not afterward. When the thermometer marks 100° C. it goes no higher, though the water boil never so hard. This temperature is called the *boiling-point* of water. On a Fahrenheit ther-

mometer this point is marked 212°. Other liquids also are purified by distillation.

Fig. 31 shows a complete apparatus, used by chemists, for the distillation of liquids. The liquid is boiled in the flask. Its vapor goes through the inside tube of the condenser *c*, while the larger tube outside it is kept full of cold water. Here the vapor is changed back to liquid

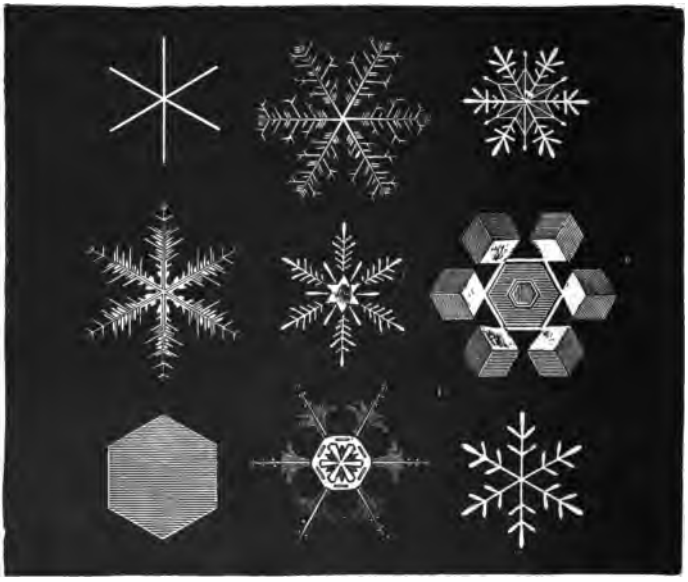


Fig. 32.

and this liquid drops into the receiver *R*. Rubber tubes carry the cold water to and from the condenser *c*. A thermometer, *t*, may be used to show the boiling-point of the liquid.

**Effects of Cold.** — A given weight of water, say 10 g., will grow smaller and smaller as we make it colder. Its

volume will be less and less until its temperature is 4° C. But if we cool it below this point the water will expand. Its volume will then become greater and greater quite regularly until it is cooled down to 0° C. At 0° the water suddenly expands much more, and at the same time it begins to freeze. This temperature at which water freezes is called the *freezing-point* of water. It is 0° C. or 32° F.

Ice is crystallized water. In blocks of ice the crystals are so crowded together that their forms are lost. The shapes of ice-crystals are seen in the beautiful frost-figures on the window-pane in winter and in snow-flakes, some of whose curious forms are shown in Fig. 32.

### EXERCISES.

**1.** *Study the effect of mixing snow and salt.*

Take snow or pounded ice, add salt gradually, stirring the two well together.

What change takes place in the two solids?

Get the temperature of the mixture.

**2.** *Compare the freezing and melting points of water.*

*1.* Put a small bottle or a test-tube partly filled with water into the freezing mixture just made, and stir this water with the bulb of a thermometer carefully until it begins to freeze.

At what temperature does the freezing begin?

Does the temperature change afterwards?

*2.* Fill a wide-mouth bottle with small pieces of ice and let it stand. Afterward stir the ice about in the water and take the temperature of the mixture.

At what temperature does the ice melt?

Repeat these two experiments and decide whether they show that there is, or is not, a difference between the melting-point of ice and the freezing-point of water.

**3.** *Find the boiling-point of alcohol, Fig. 30.*

**4.** *Find the boiling-point of a mixture of alcohol and water made in the proportion of one volume of alcohol to two volumes of water.*

Use the apparatus shown in Fig. 30.

Note the temperature when the boiling begins.

Turn the lamp low and let the boiling go on slowly until about 5 cc. of distillate is caught. Then change the test-tube.

Note the boiling-point again.

Repeat this several times, and then compare the distillates, by their odors and by means of a match-flame.

Which contains the most alcohol?

Does the liquid in the flask still contain alcohol?

The fact is that two liquids which have not the same boiling-point can be roughly separated by this process of distillation. It is called *fractional distillation.*

**5.** *Find by evaporation, whether the water in use holds any solid matter in solution.*

How, by the use of the balance and the graduated cylinder, can you find how much of this mineral substance the water contains?

## CHEMISTRY OF THE ATMOSPHERE.

Not one hundred years ago the air was thought to be an element; that it is not was proved by the great French chemist Lavoisier.

**Lavoisier's Experiment.** — The apparatus which he used was much like that shown in Fig. 33. A small quantity of pure mercury was put into a flask which was placed over a furnace. The flask had a long, slender neck, which

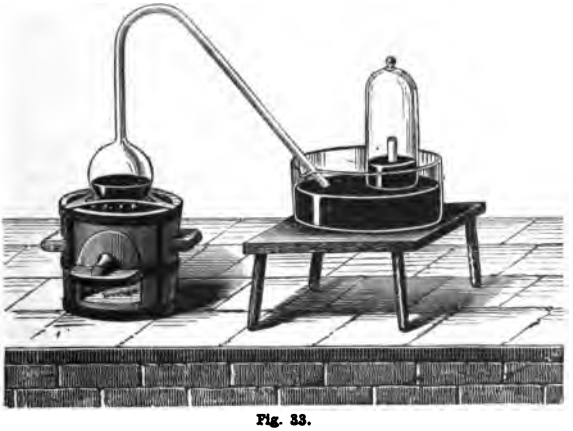


Fig. 33.

reached over into a pan of mercury. Standing mouth downward in this pan was a jar filled with air, and the neck of the flask was bent up into it.

When all was ready, Lavoisier lighted the fire in the furnace and kept it burning all the time for twelve days. On the second day he saw little red flakes of something swimming around on the surface of the mercury. For four or five days afterward the quantity of this red substance

increased while the quantity of air in the receiver diminished. For some time longer the heat was kept up, but no further change took place, and this part of the work was done. He had less air in the apparatus than at first, shown by the mercury rising in the jar, but instead of the air which was lost he had the new red substance in the flask.

What was this red substance? To find out, Lavoisier heated it in a tube so fixed that any gas which should be produced would be caught in a vessel over mercury. The red substance became black, then began to waste away while bubbles of a colorless gas were caught in the vessel prepared for the purpose, and globules of shining mercury gathered on the walls of the tube above the heated part. What was the colorless gas? Lavoisier plunged a candle-flame into it; the candle burned with a dazzling light. The gas was *oxygen.*

But whence came this oxygen to combine with the mercury when it was heated with air in Lavoisier's flask? The air must have given it to the mercury, and so the experiment proved that oxygen is one constituent of air.

In the flask and the glass jar (Fig. 33) there was still left a large quantity of air-like substance. But on plunging a candle-flame into it the flame was put out as it would have been in water. Plainly it was not air. In fact it was the gas called *nitrogen.*

Lavoisier's experiment proved that oxygen and nitrogen are two constituents of air. There are indeed a few other gases in the atmosphere beside these, but in comparison with these the quantity of them is small. Oxygen and nitrogen are the two chief constituents of the air.

### NITROGEN.

When a substance burns in air it takes the oxygen and leaves the nitrogen. Lavoisier burned mercury, but sul-

phur and some other things will burn more quickly, and may be used instead. Let us try sulphur, and afterward phosphorus.

*Ex. 43.* — I cut a slice half an inch thick from a cork which is much smaller than the mouth of my bottle. I shape the top of the cork into a shallow cup and rub it well with crayon-powder, or better with a paste of moistened plaster of Paris. I put sulphur in this cup, place the cup on the shallow water in the water-pan, set fire to the sulphur, and put a bottle bottom upward over it, as shown in Fig. 34. Describe



Fig. 34.

The flame of the sulphur.

The action of the water when the burning is over.

The change in the gas after long time standing.

*Ex. 44.* — I use a piece of phosphorus, not larger than a good-sized kernel of wheat, with another bottle holding about 200 cc. I treat it just as I did the sulphur, and again describe the flame, the action of the water afterward, and the appearance of the gas inside after standing some time over water.

But the handling of phosphorus is dangerous, unless it is done with great care. Phosphorus takes fire easily and burns the flesh cruelly. Cut it under water, lift the piece with the knife-blade, dry it by gentle contact with filter-paper, and put it into a *dry* cup. Never handle phosphorus without using the greatest care.

*Ex. 45.* — When the gas in the bottle used in Ex. 43 has become clear I slip a square of glass or of cardboard under the mouth of the bottle, lift it out of the water, turn it mouth upward, stand it on the table and leave it covered.

I at once ignite a match, uncover the bottle, and insert the flame; the nitrogen will quench it. I leave the bottle uncovered. I treat the bottle used in Ex. 44 in the same way; the nitrogen again puts out the flame. I leave this bottle, also, uncovered.

*Ex. 46.* — I now again insert a match-flame in the bottle first left uncovered, and afterward in the other. The flame is not quenched.

What does this prove?

*Ex. 47.* — I now add a little blue litmus-water to the water in the bottle in which sulphur was burned.

Note the change of color. Compare Ex. 8.

What causes this change of color?

*Ex. 48.* — I add blue litmus-water to the water in the second bottle which was left uncovered in Ex. 45; it changes from blue to red.

Can you explain this change of color?

**Burning of Sulphur.** — Sulphur, when burning with its feeble blue flame, combines with oxygen, and the two become sulphur dioxide. The water soon dissolves the whitish vapor and rises into the vessel, and at last fills just the space which the oxygen of the air occupied at first, while the nitrogen of the same air remains above the water (Ex. 43).

The sulphur dioxide shows its presence in the water by reddening the blue litmus, Ex. 47, as it did in Ex. 8.

**Burning of Phosphorus.** — When phosphorus is used the action is much the same. It combines with the oxygen of the air and forms phosphoric oxide, which fills the vessel as a milk-white vapor. Water soon dissolves this oxide, and the nitrogen of the air is left as before.

The phosphoric oxide also shows its presence in the water by reddening blue litmus (Ex. 48).

**Properties of Nitrogen.** — Nitrogen is a colorless gas (Exs. 43, 44). It is lighter than air (Ex. 46), but a liter of it weighs fourteen times as much as a liter of hydrogen. It will quench fire (Ex. 45), because it cannot unite with the elements of the fuel as oxygen does. In fact, nitrogen is the least active of the elements. It will not only quench fire, but if breathed instead of air it will quench life also. Yet it cannot be poisonous, since we inhale it with every breath without injury. It is the oxygen of the air that sustains life, and it is the absence of oxygen, and not the presence of nitrogen, which causes death when pure nitrogen is breathed.

**Other Constituents of Air.** — The air also contains water in form of invisible vapor. This is proved by placing a piece of caustic potash in an open dish. The potash will very soon become wet, and if left for some time it will be dissolved by the water which it takes from the air. Try it. The moisture to be seen on the outside of a vessel of ice-water in summer is the condensed water-vapor of the air. Dew and hoar-frost are also the water of the air, changed by cold from vapor to liquid and solid forms.

The air also contains carbon dioxide. This is shown by lime-water, which if left exposed in an open vessel will become covered in a few hours with a white crust. Try it. This crust is the same substance which is seen in lime-water after it has received carbon dioxide (Ex. 7).

The air also contains ammonia in very small quantities.

Nitrogen, oxygen, water-vapor, carbon dioxide, and ammonia are the regular constituents of the atmosphere. Our next question is, How much of each of these substances is to be found in air?

**The Analysis of Air.** — We set out now to find how many cubic centimeters of nitrogen and how many of oxygen and carbon dioxide there are in 100 cc. of air.

To do this we will imprison a vesselful of air, and then run into it a liquid which will absorb both the oxygen and the carbon dioxide completely, and leave the nitrogen. We can then measure the nitrogen which is left, and we can find out how much there was of the other two, by measuring the liquid which has gone into the tube to take their place.

*Ex. 49.* — OUR APPARATUS. — I take a test-tube, *t* (Fig. 35) to hold the air. A six-inch tube, $\frac{5}{8}$ inch in diameter, will do; an eight-inch tube of the same diameter is better. The rubber stopper, *c*, is so large that its small end will enter the tube only about a half-inch. It has two holes; to close one I have a solid rod of glass, *s*; for the other, a glass tube reaching just a very little below the cork, as shown. A piece of thin rubber tubing, *h*, is cut about six inches long. There is a pinch-cock, *p*, by which its walls may be pinched so as to close it completely. *F* is a small glass funnel.

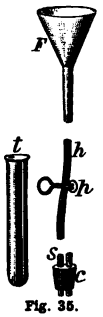


Fig. 35.

The lower end of *h* I stretch over the tube in the cork *c*, and its upper end I fix over the stem of *F*, and then I place the funnel in the clamp of the support, as shown in Fig. 36, and remove the rod *s*.

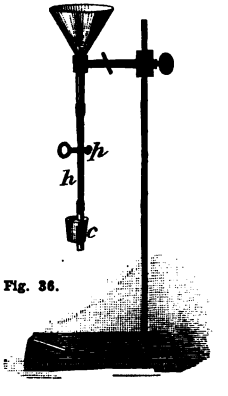


Fig. 36.

THE LIQUID. — To absorb the oxygen and carbon dioxide gases I use a mixture of pyrogallic acid and potassium hydrate.

I take a small teaspoonful of the solid acid and pour on it 10 cc. of water; it will soon dissolve. To this I then add 5 cc. of strong solution of potassium hydrate, and *at once* pour it into the funnel. Next, I hold the dish

below the cork and open the pinch-cock *p* a moment, to let the liquid run down and fill the tubes completely. I carefully take off the drop, which hangs at the lower end of the tube below the cork, with a piece of filter-paper.

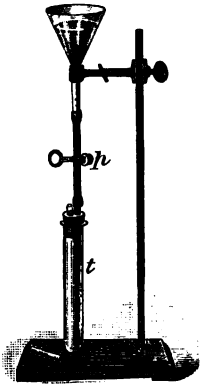


Fig. 37.

I press the tube *t* up over the cork until the joint is air-tight, as seen in Fig. 37, and after a minute I put the rod *s* into the open hole of the cork. I have now imprisoned a tubeful of air; none can get out, and no more can get in.

I left the hole in the cork open, because if it were not open the pressure of the cork would crowd the air below, and there would be too much in the tube; and then, too, handling the tube warmed it, and the volume of air changes with heat. With the hole open, the air in the tube soon comes to be just as warm and just as much pressed as the air outside. Whenever a gas of any kind is to be measured its *temperature* and *pressure* must be the *same as those of the air outside.*

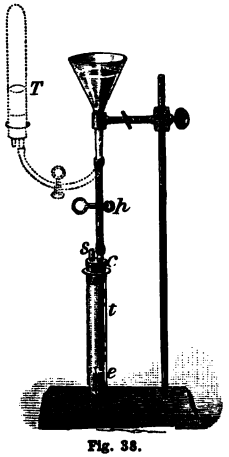


Fig. 38.

THE ABSORPTION. — I now press the pinch-cock *p;* a little stream of the liquid falls into *t* at once, and then drops follow, or, if the tube be slightly inclined, a slender stream will flow down its side. It will continue to enter as long as there is any oxygen or carbon dioxide for it to absorb, and then stop. The gas which is left in the tube is nitrogen.

But this gas is crowded down by the pressure of the liquid in the rubber tube and funnel above, and so I take hold of the cork *c*, and the rim of *t*, not to warm the gas with my hand, and lift the tube bottom up, as shown at *T* in Fig. 38, making the level of the liquid the same in the tube and in the funnel. I then open the pinch-cock. Some of the liquid will run out of *T*. When the liquid stands at the same level in the tube and in the funnel, I close the cock and bring the tube down again.

The almost black liquid in *t* has now taken out all the oxygen and carbon dioxide from the tubeful of air, and left all its nitrogen.

THE MEASURING. — I must measure the liquid in the tube to find how much oxygen was taken out,[1] and the space above it to find how much nitrogen was left.

To do this I slip two small rubber rings up on the tube, and make the upper edge of one mark the place of the lower end of the cork, and of the other, the top of the liquid. These rings must not afterward be disturbed.

I may now remove the cork, empty the tube, rinse it with water, and then let the last drop of water drain away. Finally, I use my graduated cylinder to find out exactly

How many cc. water will fill the tube to the first ring.

How many cc. water will fill the tube from the first to the second ring.

THE CALCULATIONS. — From these two numbers we can find what part of the air is nitrogen and what part is oxygen. For they help us to answer the following questions, in their order, one after another, as shown by an example below.

[1] And carbon dioxide also. But the volume of the carbon dioxide, in so small a quantity of air as we use, is so little that we cannot measure it with our apparatus. We may leave it out of account in this experiment.

How many cc. of air were in the tube at first?
How many cc. of nitrogen did this air yield?
How many cc. of oxygen did the same air yield?
Then what fractional part of the air is nitrogen?
What fractional part of the air is oxygen?
And how many cc. nitrogen in 100 cc. of air?
How many cc. of oxygen in 100 cc. of air?

*An Example.* — In an actual experiment it was found to take of

| | |
|---|---|
| Water to fill the tube to the first ring . . . . . | 6.0 cc. |
| Water to fill the tube from the first to second ring. | 23.5 cc. |
| Hence the number of cc. of air taken . . . . . | 29.5 cc. |
| And the number of cc. of nitrogen found . . . . | 23.5 cc. |
| And the number of cc. of oxygen found . . . . | 6.0 cc. |

Now this would show plainly that $\frac{23.5}{29.5}$ of the air is nitrogen and $\frac{6.0}{29.5}$ of it is oxygen. Then in 100 cc. of air there would be

| | |
|---|---|
| Nitrogen . . . . . | 79.66 cc. |
| Oxygen . . . . . | 20.34 cc. |

**The Exact Composition of Air.** — Some of the greatest chemists have devoted much time to find out exactly how much oxygen and nitrogen the air contains. They have used other methods, more accurate than that of absorption, and apparatus much more refined than that which has answered our purpose.[1] They have found the percentage composition to be

| | |
|---|---|
| Nitrogen . . . . . | 79.04 |
| Oxygen . . . . . | 20.96 |
| Air . . . . . | 100.00 |

[1] For full description, see Roscoe and Schorlemmer, Vol. I. pp. 439–447.

When the gases are weighed instead of being measured the numbers are different. Thus:

| | |
|---|---|
| Nitrogen . . . . . | 77.00 |
| Oxygen . . . . . | 23.00 |
| Air . . . . . | 100.00 |

**How much Water-vapor in Air?**—The quantity of water-vapor in the air is all the time changing. There is only a certain amount which air can hold at any given temperature, and the only way to make it able to take any more is to heat it.

But the atmosphere seldom has all that it can hold. At 60° F. the quantity will be usually found between $\frac{1}{60}$ and $\frac{1}{200}$ of the bulk of the air.

**What Fraction of the Air is Carbon dioxide?**—Sometimes as much as $\frac{1}{1000}$ of the air is carbon dioxide; at other times or places there may be no more than $\frac{1}{5000}$. Perhaps the average proportion may be $\frac{1}{2500}$.

So small a quantity as *one* cubic inch of this gas in *twenty-five hundred* cubic inches of air seems, at first thought, too small to be worthy of notice. It is not so. Carbon dioxide, little as there is of it, is one of the most important substances in the atmosphere. Plants cannot grow without it. It is a needful part of their food, and they are, when growing, all the time taking it out of the atmosphere. It is returned to the air by the breathing of animals, by the burning of fuels, by the decay of plants, and by volcanos.

We will shortly make a special study of this gas, but for the present we go on with the description of air.

**The Atmosphere a Mixture or a Compound?**—It is found that the carbon dioxide of the air makes lime-water turbid just as the same gas from marble will do. The water-vapor in the air may be condensed to liquid by cold,

just as water-vapor alone may be. Bodies burn in oxygen; they burn also in air, of which oxygen is a constituent. Nitrogen alone extinguishes fire completely, and nitrogen in the air strives to do so, but succeeds only in making them burn more feebly in the oxygen which is present. In fact, the properties of the constituents of air can be easily detected in the atmosphere itself.

But when substances combine they form a compound in which their own properties are not to be found. We therefore say that the atmosphere is not a compound, but a mixture.

**Constant and Uniform in Composition.** — The proportions of oxygen, nitrogen, and carbon dioxide are always very nearly the same. Air on the tops of mountains, and in the low places of the earth, and in different countries, have all these constituents, and very nearly the same relative quantities of them.

The heavy carbon dioxide, the lighter oxygen, and the still lighter nitrogen, are uniformly mixed together. Why do they not separate into layers, as water and oil would do after having been mixed never so perfectly?

**Diffusion of Gases.** — The fact is that gases will not stay unmixed if they touch each other. We may bring a light gas on top of a heavy one, and find that a part of the light one will sink, and a part of the heavy one will rise, until the two are completely mixed.



Fig. 39.

*Ex. 50.* — I will bring hydrogen to rest on top of air. I select two wide-mouth bottles, each holding about 200 cc., one with a neck a little smaller than that of the other, so that they may be put together, as in Fig. 39. The upper bottle is to be filled with hydrogen, the lower one with air.

To FILL THE BOTTLE WITH HYDROGEN. — I first fill it to

the brim with water; cut a piece of heavy blotting-paper a little larger than its mouth and *slide* it on as a cover and smooth it down closely upon the glass (Fig. 40, *a*). I grasp the bottle and turn it bottom upward over the pan, *B*, lower it into the water, remove the paper cover, and stand the bottle bottom upward in the water. The pressure of the atmosphere will safely hold the water in the bottle while I invert it, and while its mouth is under water in the pan.

I now put a few clippings of zinc into a side-neck flask and fix the flask in the clamp, then join one end of a

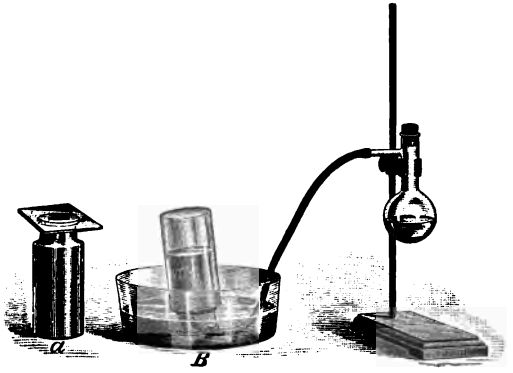



Fig. 40.

rubber tube to the side-neck, the other to a glass tube which reaches into the water of the pan. I pour some dilute hydrochloric acid (half water) into the flask and cork the neck at once. As soon as the air is driven out of the flask I put the end of the glass tube under the mouth of the bottle,[1] and collect the hydrogen as shown in Fig. 40.

[1] A little shelf for the bottle to stand on may be made by bending the ends of a strip of sheet-iron down to serve as feet to stand on the bottom of the water-pan.

When the bottle is filled with gas I again slide the cover of blotting-paper under its mouth, and hold it there while I lift this bottle and carry it over the mouth of the other. I then remove the paper, and at once let the mouth of this hydrogen-bottle down into that of the air-bottle, as shown in Fig. 39.

After waiting ten minutes I lift the two bottles, hold them over the water-pan, then separate them, and quickly put them both, mouth down, in the water.

With a lighted match in one hand, I lift the hydrogen-bottle and bring the flame to its mouth; a small explosion follows. I treat the air-bottle in the same way; there is another explosion. These explosions show that there was a mixture of hydrogen and air in each bottle.

Now air is almost 14.5 times heavier than hydrogen, and yet, to make that mixture, it must have risen into the upper bottle while the lighter hydrogen must have fallen into the lower one. They have mixed in spite of their difference in weight.

This mingling of gases, when they are simply brought into contact with one another, is called the *diffusion of gases.* The difference in weight of oxygen, nitrogen, and carbon dioxide would arrange these gases in layers one above another, the heaviest at the bottom, but the principle of diffusion will not allow this separation; it causes all these gases to mix uniformly in every part of the atmosphere.

### RESPIRATION.

Animals and plants need the gases of the atmosphere as much as they need food and soil. Without food, an animal starves; without air, its death would come still more quickly. Pluck its roots from the soil, and a plant withers and dies; give it the most fertile soil, but take away the air, and its death is just as certain.

Both animals and plants are able to take the air into their bodies and then throw it out again. This act is called *respiration.*

**Respiration of Animals.** — When an animal breathes, the air which goes to the lungs contains oxygen and nitrogen, with a little of carbon dioxide, water-vapor, and still less of a few other gases.

But what comes from the lungs? Large quantities of water-vapor, as we know by observation, as in a cold winter morning, when every breath looks like a cloud of steam. We *see* the moisture of the breath at such times, because the cold air condenses it and makes it visible.

But we can find this water, and something else, at the same time by experiment.

*Ex. 51.* — I take a clean and dry bottle, cover the lower half of its mouth with my open lips while I pass a full breath into it. Notice the deposit of dew on the walls of the bottle.

I next pour a few cubic centimeters of lime-water into the bottle and rinse the walls with it, and note the effect on the lime-water.

*Ex. 52.* — I pour 25 cc. strong lime-water into a clean bottle and add 25 cc. of water. I take a glass tube, or a straw, and, with one end in my mouth, the other in the lime-water, I breathe through the transparent liquid perhaps two or three times.

What is the effect on the lime-water?

What two substances have we found which must have been thrown out in the breath?

Now, besides the carbon dioxide and water, which we detect by these experiments, there is the nitrogen. The fact is that while the nitrogen of the air is returned to the air by the breath, the oxygen is used up in the body to make carbon dioxide and water.

How is this brought about? — The air goes into the lungs, and its oxygen passes through the pores of the delicate membrane, of which the lungs are made, *into the blood.* With the blood, as it goes from the lungs, this oxygen is carried to every part of the whole body. It is while in the blood, going from place to place, that the oxygen meets the hydrogen and the carbon.

But how came these to be in the blood? Every part of our bodies is constantly wearing out. You cannot move a finger without some of its particles being worn away. You cannot step without some of the particles of the legs being worn out. Every breath, every motion, every thought, renders useless some portions of the parts of the body which are acting at the time.

These worn-out particles are all the time going into the blood. They consist chiefly of hydrogen and carbon, and it is from these worn-out particles of our bodies that the oxygen gets hydrogen and carbon to form the water-vapor and the carbon dioxide which are thrown out at every breath.

Life itself depends upon this process. These waste particles are the impurities of the blood, and they must be taken out or death will quickly come. As long as oxygen is regularly supplied to the lungs it will change these impurities into water and carbon dioxide, and in these forms they will be thrown out in the breath.

These two seem to be the most abundant substances thrown out, but they are by no means the only ones. Many other impurities are exhaled at the same time. Some of these are very offensive, and all of them are very injurious to health if taken again into the lungs.

**Air spoiled by breathing.** — Oxygen is the only thing in the air that can purify the blood, and this element is being taken from it by every breath. Once breathed, the air is unfit, on this account, to be breathed again.

This is only one way, and that not the worst, in which the act of breathing spoils the air. Every breath is polluting all the air into which it is thrown, with the impurities of the blood. Air without its oxygen would not be poisonous nor filthy, but air, with the abundant impurities thrown into it with the breath, is both poisonous and filthy. Were we able to see the impurities of air which has been breathed, we would shun such atmosphere as we now shun the water of a stagnant pool. Ought we to shun it less because they are invisible?

**Ventilation.** — It is necessary to change the air in the rooms of our houses very often and very thoroughly, in order to avoid being poisoned by our own breath.

The removal of foul air and the introduction of that which is pure is called *ventilation*. Every room in which human beings are expected to live ought to have some means of ventilation.

Our chemistry teaches us these facts, but chemistry is not left to do this alone. The same lesson is taught by some of the most awful experiences.

Sometime more than a hundred years ago it happened that, in Calcutta, one hundred and forty-six persons were shut up for a night in a small room called the Black Hole. At dawn of day only twenty-three remained alive.

The passengers on board a ship were all crowded into the cabin, one stormy night. One hundred and fifty went in, but only eighty came out alive.

Examples of less painful kind are much more common. A schoolroom is unventilated; the pupils become listless and dull from the influence of bad air. A church or hall is not ventilated, and a large audience becomes languid and sleepy. A bedroom has its doors and windows tightly shut; the sleeper awakes in the morning unrested and with headache.

These are the effects of breathing air which has already been breathed.

Every building in which people are to live, even for short times, ought to have the means of ventilation "built in," but if the builder of the house has not provided some special means, then the windows and doors should be freely used for the purpose. There should be two openings in every inhabited room, one near the top, the other near the bottom. Out of one of these the foul air may escape while fresh air may come into the room through the other. A window let down a little from the top, and raised a little from the bottom, will meet this condition in a degree.

**The Respiration of Plants.** — Plants breathe. They take air into their leaves and throw it back into the atmosphere again; and this is respiration.

The leaves are curiously made. Look at one with a good microscope, and we see that its surface is covered with little openings, or *pores*. What a multitude of them! In some cases more than a hundred thousand in the small space of one square inch! They are found on both the under and the upper side of the leaf. The air is taken into these little mouths on the *under side*, while from those on the *upper side* the gases are thrown out.

Animals breathe in order to get oxygen from the air; plants breathe in order to get water-vapor and carbon dioxide. The carbon dioxide is decomposed; its carbon is kept in the plant while its *oxygen is thrown out*. While the sun shines this action goes on; in the dark it does not, for it is found that, at night, carbon dioxide is thrown out instead of oxygen.

Animals are all the time using up the oxygen of the atmosphere; plants are by the help of sunlight throwing it into the atmosphere. Animals are constantly giving water-vapor and carbon dioxide to the atmosphere; plants

are busy in taking these substances out. Thus do these two great kingdoms of nature balance and sustain each other.

### EXERCISES.

**1.** *Study the action of sulphuric acid on oxalic acid.*

*1.* Bring small quantities of the two substances together in a test-tube, using the concentrated sulphuric acid, and crystals of oxalic acid. Discover whether chemical action will take place in the cold, the effect of heat, whether a gas is set free, and any other facts you can. This work will suggest the next step.

*2.* Proceed to collect some of the gas in order to learn what it is. To do this, put 5 g. of oxalic acid, in crystals, with 15 cc. of strong sulphuric acid into a side-neck flask, to make the gas; and to collect it, arrange two conical flasks and one test-tube, which is fitted with a cork and tubes in the same manner as the flasks.

*3.* Examine the gas. Put a flame into the mouth of the first flask to learn whether the gas is combustible or otherwise; and for the name of the gas which gives the result you get here consult p. 46.

Test the gas in the second flask with lime-water, and name the gas which this test reveals.

Having now proved that the gas which is set free by the action of the two acids is a mixture of carbon monoxide and carbon dioxide, you can take the fourth step in your investigation, which is to

*4.* Analyze the mixture. Do this by the method of "absorption," the same as that used in the analysis of air in Ex. 49.

Fill the funnel and the rubber tube, shown in Fig. 36, not using the cork, with a strong solution of potassium hydrate. This liquid will absorb carbon dioxide, but not

carbon monoxide. Then slip the end of the rubber over the end of the long glass tube which passes into the test-tube in which you caught the gas. Put a solid stopper in place of the short glass tube, and then go on with the work as in Ex. 49. When you come to measure the volumes do not forget that you have used the long glass tube in the test-tube full of gas.

*5.* Finally, write a brief account of your investigation and its results.

**2.** *Study the action of phosphorous on air.*

We have seen that phosphorus burns in air when it is heated (Ex. 44); but how will the two behave if no heat is used? Remember to handle phosphorus with great care. See p. 67.

*1.* Place a small piece of phosphorus on the end of a small wire, and put it into a bottle. Then invert the bottle and leave it standing, mouth downward, in a vessel of water several hours. Notice any evidence of chemical change or of absorption, or of any other action which may occur.

*2.* Remove the bottle from the water without losing any part of what is in it, and at once test the gas which is left. Decide what this gas is, and then what the phosphorus must have done.

*3.* See if you can find out, by measuring, what fraction of the air the phosphorus took out.

*4.* Write a brief account of your work and results.

## COMPOUNDS OF NITROGEN, HYDROGEN, AND OXYGEN.

We have seen that about four-fifths of the air is nitrogen gas, and that its effect is simply to dilute the oxygen so that combustion will be less furious than it would be in oxygen alone, and so that the air may be mild enough for animals to breathe. Nitrogen is just fitted to do this, because it has so little affinity for other elements.

But nitrogen can be made to combine with other elements, and in fact there are a great many of its compounds known. Many of these are strangely unlike the element itself, for while it is the mildest of all things, these compounds are among the most pungent and corrosive substances to be found.

### AMMONIA.

Ammonia is the compound of nitrogen and hydrogen.

**Produced in two Ways.** — There are two ways in which these two elements can be made to combine: one is by electricity. If we mix the gases, and then send a silent electric charge through the mixture for a long time, we find some ammonia in the tube.

Ammonia in small quantities is found in the air, and it is likely that the electricity of the atmosphere has produced it there.

The other way to make nitrogen and hydrogen combine is to bring them together just at the moment when they are set free from some other compounds. For example, they are both found in animal bodies, such as horns, hoofs, and flesh. Now, when these decay, the nitrogen and hydrogen are set free, but they are together at the same moment,

and combine with one another. Neither escapes alone, but they go off together as ammonia. Ammonia may be found in the air of stables, of farm-yards, and of other places where animal matters are decomposing.

THE NASCENT STATE. — These gases are not the only things which combine most readily just at the moment when they are set free from something else. Other elements also seem to be more active at that moment than at other times. They are then said to be in the *nascent state.*

FORMED IN GAS-WORKS. — Gas for lighting towns and cities is made by heating soft-coal to a cherry-red heat. Among the gases and vapors driven off are ammonia and some of its compounds. These are bad impurities, and in the manufacture of gas they are washed out by cold water. The ammonia, in this "ammonia water" of the gas-house, is saved by making it combine with hydrochloric acid, which changes it to *ammonium chloride,* or with sulphuric acid, which changes it to *ammonium sulphate.* This is the source of nearly all the ammonia and its compounds, great quantities of which are used annually.

**Preparation of Ammonia.** — Ammonia itself is made by decomposing one of these compounds by heating it with lime.



Fig. 41.

*Ex. 53.* — I powder a very little ammonium chloride and also a little good quick-lime. I then mix them well, and put into a clean and dry test-tube enough to half fill the rounded bottom. I fill the mortar with water colored blue by litmus, and add only a drop of hydrochloric acid to redden it: let this stand ready for use when needed. I take the tube between my fingers, with my thumb over

its mouth, leaving a small opening at the lower edge, and hold it some time in the hot air above the flame, a little inclined, as shown in Fig. 41. Note

The deposits of moisture on the cold parts of the tube.
The odor at the mouth of the tube.
This is the odor of ammonia.
What two substances seem to have been produced?

By this time the decomposition of the mixture is at an end, and I bring the mouth of the tube down into the reddened litmus-water, still holding the tube so much inclined (Fig. 42) that the white solid will not fall down into the water. As the tube cools the red water rises in it. But it will rise much further than it would if the tube were heated without the mixture. (Try an empty tube in the same way, and see how far the water will rise.)



Fig. 42.

What property of the gas does this show?
Note the change of color in the water.
Compare this change with that in Ex. 48.

THE FACTS. — We find that by heating the ammonium chloride with lime we get *water*, condensing in droplets on the tube; *ammonia*, which we know by its familiar pungent odor; and a white solid left clinging to the bottom of the tube, which is *calcium chloride*.

We see, too, that ammonia is a gas, but a gas which is soluble in water. Now the so-called ammonia in commerce is a liquid. It is, clearly, not the real ammonia. The fact is, it is water with as much as it will hold of the ammonia gas in solution.

One cubic centimeter of water at 15° C. will hold 783 cc. of this gas, and, like all other gases, ammonia is more soluble in colder water. At 0° C. the cubic centimeter of

water will absorb 1148 cc. of this gas. This strong solution of ammonia is known as *ammonium hydrate.*

By warming this liquid ammonia, the gas itself can be obtained in great abundance.

*Ex. 54.* — To do this I arrange the apparatus as in Fig. 43. I put 10 cc. or 15 cc. of the liquid ammonia in the side-neck flask, close its mouth and connect it with the *short tube* of *a*, whose long tube is joined to the short one of *b*. The stoppers must close the flasks air-tight.

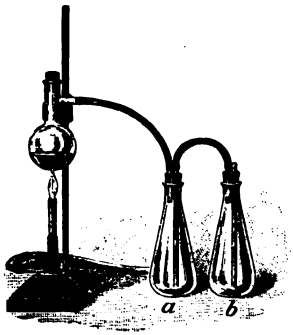


Fig. 43.

I put the *short tube* of *a* next the flask, because this gas is lighter than air. It will collect in the upper part of *a*, and push the air out through the long tube into the upper part of *b*, and afterwards out of the long tube of *b*, until both flasks are full.

I now make the lamp-flame very small, so that only a current of hot air will warm the liquid in the flask. The gas will soon escape freely. Some water-vapor goes over with it, but I use as gentle heat as I can to keep it bubbling; the water does not boil; the bubbles are ammonia gas, mixed with little vapor.

Note the odor of the gas which escapes from the open tube of *b;* it will be more and more pungent until the air is all out of the flask.

*Ex. 55.* — Moisten a stick with hydrochloric acid and hold it just above the tube of *b*.

What evidence of chemical action do you find?

Compare this with Ex. 9.

*Ex. 56.* — I take both rubber tubes from *a*, and at once plunge the stoppered mouth of the flask down into a vessel of water (Fig. 44).

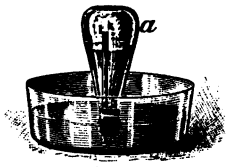


Fig. 44.

What is the result?

What does it teach about ammonia gas?

*Ex. 57.* — I redden some blue litmus with as little hydrochloric acid as will do it, and then pour into it some of the water which has just now dissolved the gas in *a*.

How does the ammonia show its presence in the water?

Has this ammonia-water any taste or odor?

THE FACTS. — Dry ammonia-gas is invisible and much lighter than air. Its solution in water has a caustic taste, and will bring back the blue color of litmus which has been reddened by acid.

Ammonia is a stimulating gas to breathe: on this account it is used to revive the faint, and sometimes to overcome the effects of an anæsthetic, like ether or chloroform.

Ammonia and hydrochloric acid combine at once, whenever brought together, and form ammonium chloride. The white cloud in Ex. 55 was a cloud of ammonium chloride.

**Ammonium Salts.** — Ammonia acts also on other acids, and so forms a large number of compounds, which as a class are called *ammonium salts*.

*Ex. 58.* — I put 10 cc. of water into a small wide-mouth bottle and add 2 cc. strong sulphuric acid. I place this in a porcelain dish and drop into it a small clipping of blue litmus-paper.[1] I now add strong ammonia water (ammonium hydrate) little by little, shaking the mixture well

[1] Filter-paper soaked in a strong solution of litmus and afterward dried.

each time, until the red paper turns blue. If now I add drop by drop of the dilute acid, I can reach the point when the paper is just reddened by the last drop, and when another drop of ammonia will make it blue again. I do this. And then I put the dish over a small flame and evaporate (Fig. 10) the liquid, until a white rim is seen around its edge on the dish, and then let it cool.

While waiting for the liquid to evaporate I begin the next experiment.

*Ex. 59.* — I mix 5 cc. of water and 5 cc. strong nitric acid, add the blue litmus-paper and then the ammonium hydrate, until the last drop turns the paper blue. I then evaporate this liquid to about one-fourth its bulk, and let it cool.

If the liquids were evaporated enough, crystals of a white solid will be seen in each when it is cold, or the whole may become a solid white mass.

*Ex 60.* — I take a little of one of these white solids to test its action on litmus. I first pour upon it, and then quickly off again, a little water — just to wash it. And then I add water to dissolve it. I take two clippings of blue litmus-paper, moisten one with water, and then redden it by holding it in the mouth of the bottle of hydrochloric acid, and then put both into the solution. No change of color should occur in either.

These white solids, made by the action of ammonia on sulphuric and nitric acids, are examples of ammonium salts.

But why did we use the litmus? Simply to show when the chemical change was ended. We have found that blue litmus is reddened by an acid in every case, and that ammonia will bring back the blue color. But when there is just enough ammonia to use up all the acid,

neither the blue nor the red is changed. The ammonia and the acid *neutralize* each other, and the litmus shows the end of this action.

Do not leave this until you see clearly *why* the two neutralize each other. Consult p. 57 for the law.

**Composition of Ammonia by Volume.** — It is found by analysis that ammonia always yields just three times as many cubic centimeters of hydrogen as of nitrogen. Its composition by volume is:

3 measures hydrogen to 1 measure nitrogen.

But another question is, How much ammonia gas would these 3 of hydrogen and 1 of nitrogen make? It has been found by analysis that 20 cc. of ammonia will give 30 cc. of hydrogen and 10 cc. of nitrogen. That is to say, 30 cc. of hydrogen and 10 cc. of nitrogen make 20 cc. of ammonia.

Three measures of hydrogen and one measure of nitrogen make two measures of ammonia. Or *four* volumes of the constituents are condensed to *two* volumes of the compound.

Compare this curious fact with another noted under the composition of water, p. 56.

## NITRIC ACID.

Nitric acid is a compound of nitrogen, hydrogen, and oxygen. It is one of the strongest of acids, and one of the most useful.

**Occurs in Combination.** — Very small quantities of this acid exist, free, in the air: it is in its compounds that it is mostly found. Two of these compounds occur in large quantities in some parts of the earth: they are "saltpetre," whose true name is potassium nitrate, and "Chilian saltpetre," or sodium nitrate.

**Made from Sodium Nitrate.** — To get the nitric acid the sodium nitrate is heated with sulphuric acid. Sulphuric acid is still stronger than nitric acid, and will drive the latter out of the nitrate in order to take its place. The nitric acid is driven out in the form of vapor, but it is carried into a cold vessel, where the vapor is condensed to a liquid.

**Properties.** — Nitric acid when pure has no color, but it is intensely sour and very corrosive. It will redden litmus like other acids, but will not stop with this: it will go right on to destroy the red color. Try this by putting a piece of litmus-paper into a little of the strong acid. It will first turn red, soon afterward yellow, and in a little time the paper will easily break to pieces. In the same way it will form yellow stains upon the fingers, and upon the garments, and, in fact, upon all organic bodies. The acid may be washed away by water, or it may be neutralized by ammonia, but the yellow color will remain.

*Ex. 61.* — Wet some pieces of woolen cloth with drops of nitric acid, with drops of hydrochloric acid, and of sulphuric acid. Wait until the color is changed by all, and then try the effect of ammonia on them. Note the difference.

The strong nitric acid in commerce is by no means pure: in fact, it is almost half water, and there are several other impurities besides water, in small quantities. By distilling a mixture of this acid with strong sulphuric acid the pure and concentrated nitric acid is obtained.

**Decomposition of Nitric Acid.** — This acid is very easily decomposed. It is decomposed by light. When standing in the sunlight, the upper part of the bottle of acid will, after a time, be seen full of a reddish vapor; this is set free from the acid by light, and water is formed at the same time. The yellow color of the com-

mon acid is due to this reddish vapor, some of which is dissolved in the liquid.

THE NITRATES. — Nitric acid is easily decomposed by almost any metal. One can see this action by putting small bits of copper or lead in a test-tube and covering them with dilute nitric acid. But the fumes are very noxious, and the experiment should be made on a small scale, or else in the open air. The same red vapor, seen when the acid is decomposed by light, is set free in abundance by the metal. Water is also formed, and beside these a *salt* of nitric acid is produced. These salts of nitric acid are called *nitrates*. The salt made with copper is blue, and it is called copper nitrate; that made with lead is white, and is called lead nitrate.

This decomposition of nitric acid by a metal will be better understood further on, when we study Ex. 62.

These nitrates are all soluble in water, and so when a metal is changed into a nitrate the metal seems to dissolve, but it is always the nitrate which is in solution, and not the metal.

AQUA REGIA. — Nearly all the metals are thus changed into nitrates by nitric acid; but gold and platinum are the exceptions. To dissolve these metals we need a mixture of nitric and hydrochloric acids. Neither of these alone can attack gold or platinum, but the two together will dissolve them readily. The mixture of these two acids is called *aqua regia*. But, as we shall see, the metals dissolved in aqua regia are not changed into nitrates, but into compounds of chlorine, called *chlorides*.

### NITROGEN OXIDES.

**Investigate the Decomposition of Nitric Acid.** — We have just seen that when nitric acid is decomposed reddish vapors are set free. What are these vapors?

Let us study the action by experiment. We will decompose the acid by copper.

*Ex. 62.* — I fit up my apparatus for making and collecting gases as usual for a gas heavier than air (Fig. 45), the long tubes of the flasks *a*, *b*, *c* toward the side-neck flask. Into both *a* and *b* I put water, which the cut shows only in *a*, but none in *c*. After the connections are made, I put about seven grams of small pieces of copper-foil, or thin sheet-copper, in the side-neck flask, pour in about 25 cc. of dilute nitric acid, half water, and close the flask with its air-tight stopper. The chemical action sets in at once. No heat is needed to start it.

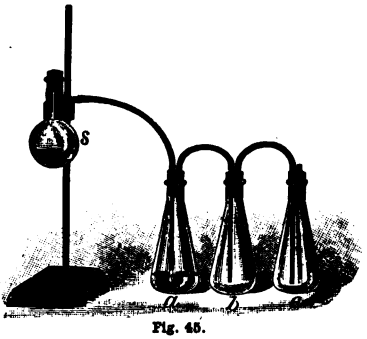


Fig. 45.

Note the change in the color of the acid.

Trace the changes of color as they occur in *s*, *a*, *b*, *c*.

We see that a red gas quickly fills the flask *s*, and goes over into *a*, *b*, *c*, but that the red color disappears from *s*, *a*, and *b*, and stays in *c* for a much longer time.

How many kinds of gas are evidently in the flasks?

When *a* and *b* are clear, and *c* still full of red-brown gas, I disjoin the rubber tube of *c*, remove the stopper carefully, pour in gently a few cubic centimeters of water, and return the stopper to its place as tightly fitted as before. I now close the ends of both glass tubes with my finger, which I can easily do if they are at the same level, lift the flask and shake it well. Note the disappearance of the red gas. I next plunge the finger and tubes under water, and take the finger away. Note the inrush of water.

Why did the color of the gas disappear with water?

Why did the water rush in when the tubes were opened?

What property of the red gas is thus discovered?

*Ex. 63.* — I now disjoin the rubber tubes of *b*, and then invert the flask and let the water run out of the short tube. Of course air will run in through the long tube and mix with the colorless gas within. The red gas is reproduced!

This is an important discovery. If air will change the colorless gas into the red gas, perhaps it was in this way that the red gas was made in the flasks at first, for they were full of air when the action began in *s*. If I could keep the air away, while the copper and nitric acid act on each other, would the red gas appear at all?

Now I can keep the air away, and answer this question, by putting carbon dioxide into the flasks in place of the air.

*Ex. 64.* — I disjoin the rubber tubes of *a*, and empty the contents of *s* into a porcelain dish to be examined afterward. I then join the side-neck with the long tube of another flask (Fig. 46). I put 20 cc. dilute nitric acid, half water, into *s*, drop in gently two or three pieces of marble as large as peas, and close the flask. Carbon dioxide will be set free, drive the air before it out at the open short tube, and will soon fill the flask. I then open *s*, drop in some bits of copper, and quickly close it again. The nitric acid and copper now act in the absence of air. Do they yield the red gas? None will be seen if the air was all expelled. Do they yield the same colorless gas as before? I mix air with the gas in *c*; the red gas instantly appears, as it did in Ex. 63.

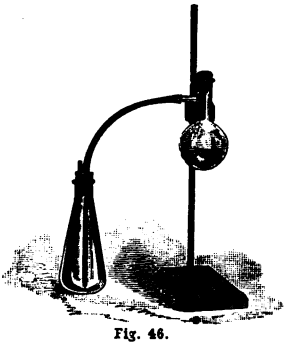


Fig. 46.

Our experiments clearly prove that when copper decomposes nitric acid, only a colorless gas is set free. But when this colorless gas meets with air, another, a red-brown gas, is formed. At the same time a blue compound is made which stays in the liquid. The colorless gas is a compound of nitrogen and oxygen, called *nitric oxide*, and the red-brown gas is another compound of nitrogen and oxygen, called *nitrogen peroxide*. The blue compound is *copper nitrate*. It may be obtained in blue crystals by filtering the liquid and evaporating it until, on cooling, the crystals form. (Try it.)

When nitric oxide mixes with air it instantly combines with oxygen and becomes nitrogen peroxide.

**Nitrous Oxide.** — There is another compound of nitrogen and oxygen called *nitrous oxide*. It is not usually made by the decomposition of nitric acid directly, but by decomposition of ammonium nitrate, and this is done by means of a gentle heat, which breaks the nitrate into water and nitrous oxide. Thus:

*Ex. 65.* — I put from 7 to 10 grams of ammonium nitrate into the side-neck flask, which should be dry, and join the flasks *a*, *b*, *c*, as usual for a heavy gas. To condense the steam, I put the empty flask *a* into a dish of cold water, — ice-water is best. It may be held down in the water by a cord passing over the stopper and around under the heavy dish of water. To decompose the nitrate I use the gentle heat of a small

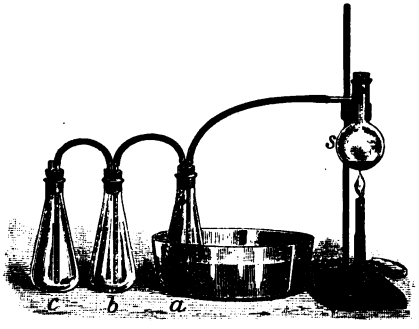


Fig. 47.

flame, just hot enough to melt and keep it gently boiling. The nitrate slowly wastes away, and when but about one-fifth remains I withdraw the lamp.

Flasks *s* and *a* prove that water is produced. How?

Test the gas in *b* with the flame of a splinter of wood.

Note the color, and the odor, of the gas in *c*.

THE FACTS. — Ammonium nitrate contains nitrogen, hydrogen, and oxygen, and when it is decomposed the nitrogen and hydrogen separate, each taking a part of the oxygen with it. The nitrogen and oxygen form the nitrous oxide, while the hydrogen and oxygen form water. The water is condensed in *a* (Fig. 47), while the nitrous oxide is collected in *a*, *b*, and *c*.

Nitrous oxide is a colorless gas in which bodies will burn with almost the same vigor as in oxygen itself. But its most remarkable action is upon a person who breathes it. Breathed in small quantities it intoxicates, and often causes a disposition to laughter. On this account it is commonly known as *laughing-gas*. But if its use be continued a few minutes, it will produce complete insensibility, and if continued, death. It is much used to render patients insensible to pain in minor surgical operations, such as the extraction of teeth. The insensibility lasts a few minutes only; it is quickly banished by fresh air.

**The Five Nitrogen Oxides.** — There are two other compounds of nitrogen, making in all five. Here are five entirely different kinds of substance made out of the same two elements. But how can the same elements make different compounds? By combining in different proportions. When the colorless nitric oxide touches the air it takes in *more oxygen* and becomes the red nitrogen peroxide. Different weights of these elements unite.

Let N stand for nitrogen and O for oxygen. By analysis it has been found that

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| Nitrous oxide | contains | 28 | of | N | and | 16 | of | O |
| Nitric oxide | " | 14 | " | " | " | 16 | " | " |
| Nitrogen peroxide | " | 28 | " | " | " | 48 | " | " |
| Nitrous anhydride | " | 14 | " | " | " | 32 | " | " |
| Nitric anhydride | " | 28 | " | " | " | 80 | " | " |

The properties of a compound depend on the *relative weights* of the elements in it. This fact is clearly shown by these nitrogen oxides.

**The Law of Multiple Proportions.** — It is a curious fact that the relative quantities of oxygen in this table of nitrogen oxides are all either 16, or else 2 or 3 or 5 times 16. So, too, in the case of the nitrogen the quantities are either 14, or twice 14. In both cases alike, the larger numbers are all exact multiples of the smallest. Now chemists have found this to be true in a great many other cases where two elements make more than one compound. In fact they have met no exception. This causes them to feel quite sure that all elements are alike in this respect, and they state this conclusion as follows:

*If one element combines with another in more than one proportion, these proportions are all exact multiples of some one fixed number.*

And this important statement of fact is known as the "law of multiple proportions." The student should now remember the "law of constant proportions" (p. 57), for these two laws together cover the most vital facts about the combination of elements.

**Combination of Definite Weights.** — These two laws show very clearly that the elements never combine except in certain definite weights.

The definite weights of oxygen and hydrogen in water are in the ratio of 16 to 2. Sixteen grams of oxygen and two grams of hydrogen will unite without leaving any of either. But if we pass an electric spark through a mixture

of 16 g. of oxygen with 3 g. of hydrogen, just 1 g. of hydrogen will be left.

The definite weights in which oxygen and nitrogen always combine have the ratio of 16 to 14.

The definite weights of hydrogen and chlorine which make their compound, hydrochloric acid, are as 1 to 35.5. In other compounds of chlorine the definite weight of chlorine used may be 2 or 3 or some other whole number of times 35.5. And in other compounds of hydrogen the definite weight of hydrogen may be 2 or 3 or some other whole number of times 1.

**Combining Weights.** — These numbers which show the ratios of the *smallest* definite proportions by weight in which these elements ever combine with anything else, are called *combining weights.* Thus we say

| | | | | | | |
|---|---|---|---|---|---|---|
| The combining weight of | | | | Hydrogen | is | 1 |
| " | " | " | " | Oxygen | " | 16 |
| " | " | " | " | Nitrogen | " | 14 |
| " | " | " | " | Chlorine | " | 35.5 |

And we mean that the smallest weight, of chlorine for example, which we can find in any compound is just 35.5 times as large as the smallest weight of hydrogen which can be found in any of its compounds.

To find these numbers — the combining weights — is one of the most difficult problems in chemistry. The student is not yet ready to see how it is done. But it has been done for the whole list of elements. Every element has a combining weight assigned to it.

### *EXERCISES.*

**1.** *Make some Nessler's reagent as follows:*

*1.* Make a solution of potassium iodide, say 6 g. in 20 cc. of distilled water, or of the best spring water.

*2.* Make a strong solution of mercuric chloride, say 6 g. in 60 cc. of water.

*3.* Add the mercuric chloride solution, little by little, to the potassium iodide, until a small portion of the precipitate will not disappear when shaken or stirred.

*4.* Make a strong solution of potassium hydrate, say 10 g. in 15 cc. of water, and when cold add this to the mixture already made.

Finally, let this mixture stand until it is perfectly clear. The clear liquid is called *Nessler's reagent.*

**2.** *Study the effect of Nessler's reagent on solutions containing ammonia.*

*1.* Add a drop or two of ammonium hydrate to a test-tube nearly filled with distilled or spring water. Then add 1 cc. of the Nessler's solution. Mix it well, and note the *color* of the mixture.

Find out whether less than a drop of ammonia will yield this color in the same quantity of water. Or how little you can use and still be able to detect the color.

*2.* Put a little solution of ammonium chloride in the water, instead of ammonia, and test it with the Nessler's solution.

Does the ammonia of the ammonium chloride yield the same color?

The fact is that Nessler's solution is the most delicate test for ammonia, always showing the presence of that substance by a yellow color, which has a light straw tint when very little ammonia is present, and a deep orange tint when there is much.

*3.* Learn by this test whether rain-water contains ammonia.

Why should rain-water contain ammonia?

*4.* Test the drinking-water of the neighborhood for ammonia.

**3.** *Study the effect of nitric acid on ferrous sulphate, sometimes called "copperas."*

*1.* Into a little dilute nitric acid in a test-tube drop a good crystal of the ferrous sulphate. Do not shake it. Notice the *color* which appears in the liquid around the crystal.

*2.* See how dilute you can make the acid and still be able to detect this color when a crystal of the sulphate is used.

*3.* Try the solution of a nitrate, say potassium nitrate, instead of nitric acid, in the same way.

Does the color now appear around the crystal? If not, there is probably little or no free nitric acid present.

*4.* Mix a little solution of the potassium nitrate with a solution of the sulphate, incline the tube, and let a little concentrated sulphuric acid run down the inside of the glass, to the bottom. Do not shake it.

The color should now appear as a ring, where the liquids touch each other.

If so, then why did it not in the other case?

The fact is that strong sulphuric acid decomposes the nitrate and sets its nitric acid free. And then the *free* nitric acid shows itself by the color it makes by acting on the crystal.

*5.* Make the experiment again with potassium nitrate and sulphuric acid, but add the ferrous sulphate to the liquid when hot.

Ferrous sulphate is a delicate and much-used test for nitric acid and the nitrates. It must be used in *cold* solutions: the brown color disappears on heating.

*6.* Get a little of some white solid, the name of which you do not know, from the teacher, or a friend who knows what it is, and see if you can tell by the copperas test whether it is a nitrate or not.

## THE COMPOSITION OF PLANTS.

THE substances which exist in plants, or which may be made from them, are so many that we cannot undertake to study them all. The most that we can do now is to ascertain the *elements* of which plants are made, and perhaps a few of their simpler compounds.

Let us examine a piece of wood as follows:

*Ex. 66.* — I take a splinter of well-dried pine wood, about as large and long as a common match, drop it into a test-tube, and then heat it slowly by holding the tube almost horizontally just above the tip of the lamp-flame, and move the tube back and forth to heat the length of the wood.

Notice the dew on the cold walls; what does this show?

Notice the vapors; do they condense or go off as gases?

Notice whether any other liquid than water is formed.

Describe the stick after the changes are over.

**The Facts.** — When wood is heated in a close vessel, where there is not air enough to burn it, there are many new products of its decomposition. Some are gaseous; these escape into the air. Some are liquid; these condense on the cold parts of the vessel. A black solid is left which no heat seems to affect. Among the liquid products is water, — the dew seen on the walls of our tube; and the black solid which was left behind is charcoal, which is the element *carbon*, almost pure. But water is always made up of *hydrogen* and *oxygen*, and as water comes from the well-dried wood, the wood must contain these elements also. Thus we prove that wood contains carbon, hydrogen, and oxygen.

By extending our experiments to other kinds of vegetable matter we find in them the same elements: carbon, hydrogen, and oxygen. Every blade of grass, every leaf and flower, and every kind of seed and fruit, contain these three elements, and they contain very little of any others.

Plants do indeed contain other elements than these. *Nitrogen* is found in them; in small quantities to be sure, but it must not be overlooked, because it is one of the most important elements in the vegetable food of animals.

Besides nitrogen there are several other elements in minute quantities, which, with carbon, hydrogen, and oxygen, enter into the composition of plants.

When wood or any other vegetable matter is burned, these four elements disappear, and nothing but a little ash remains; but this ash contains all the other elements of the substance. How small the quantity! It seldom amounts to one tenth of the whole. In every 100 pounds of vegetable matter, from 90 to 97 pounds are made up of carbon, hydrogen, oxygen, and nitrogen. All these facts have been established by experiments.

**The Food of Plants.** — An animal, to live and grow, must be supplied with food. The same is true of plants: they must be supplied with nourishment which contains all the elements they need to promote their growth. Now carbon, hydrogen, and oxygen, with a little nitrogen, and very small quantities of a few other elements, make up every part of any plant. These the plant must get in some way, else it cannot flourish. If any of them are lacking, the plant suffers even if it does not die. The food of plants must contain all these elements.

The plant gets its carbon from *carbon dioxide*. It gets its hydrogen and oxygen from *water*. It gets its nitrogen from *ammonia*, and as nine-tenths of the weight of plants is made up of carbon, hydrogen, oxygen, and nitro-

gen, it is easy to see that the three substances just named must be the food upon which plants must chiefly live.

BUT HOW CAN A PLANT TAKE FOOD? — Every plant has a multitude of mouths. There is one at the end of every little rootlet in the soil, and there are a host of them on the *under side* of every leaf. Each little root-mouth of a growing plant is taking in *liquid* food from the *soil*, and each little leaf-mouth is at the same time taking *gaseous* food from the *air*.

The liquid which enters the roots of a plant is water in which many substances of the soil are dissolved. It contains compounds of ammonia, from which the plant can get nitrogen, and it also furnishes the plant with those other elements which it needs in small quantities, those which make up the *ash* which is left when the plant is burned.

The gaseous food which enters the leaves of a plant is the carbon dioxide and water-vapor of the air. From these gases the plant gets carbon, hydrogen, and oxygen, — the three most abundant elements needed in its growth.

The food of plants must be *decomposed* before its carbon, hydrogen, oxygen, and other elements can nourish them. These elements then combine again in very different ways to produce all the materials of which the parts of a plant are made, such as starch, sugar, wood-fiber, gums, oils, and coloring matters.

Among the four chief elements in plants, the only one which we have not studied is:

### CARBON.

**The Source of Carbon in Plants.** — Carbon dioxide is one of the most important substances in the food of plants. But the oxygen of this substance is of no use to them; they get enough of that from other sources;

it is the carbon which is useful. Now, every leaf of a growing plant is taking the carbon dioxide of the air which passes over it. We learned this when studying the subject of respiration. This substance is decomposed while in the leaf; its oxygen is exhaled, but its carbon remains to enter into combination as a part of the body of the plant.

**Charcoal-making.** — This carbon, which the plant has taken from the atmosphere, and also the little which it may have taken up through its roots, is obtained without



Fig. 48.

much difficulty, for use in the arts, and we are acquainted with it under the name of charcoal.

The charcoal-maker piles his sticks of wood in the form of a mound, and covers the whole with dirt and turf. He leaves a few small holes for a little air to enter the pile

at the bottom, and another at the top for the smoke to escape, and in this way a half-smothered burning is kept up for a long time, as shown in Fig. 48.

Now what change occurs? A very simple one. The wood is decomposed by the heat; its gaseous constituents are driven away, but its solid carbon is left behind. Not all of it, to be sure, for a little of it unites with oxygen and flies away as carbon dioxide. But the carbon which is lost in this way is very little compared with the charcoal which is left.

In appearance, the wood only seems to have changed in color. It is black. In other things the charcoal looks like the wood. There is the bark with all its knotty roughness. There are the annual rings inside the bark, to be plainly counted, and, if we look through a microscope, there are the delicate cells, which the microscope could have shown us in the wood before it was burned. Let us lift it, and it is easy to feel that the wood has lost much of its weight, but really it is not easy to see that it has lost much of its size; the stick of charcoal is very nearly as large as the stick of wood from which it was made.

CHARCOAL IS NOT QUITE PURE CARBON. — When pure carbon burns it is wholly changed into carbon dioxide gas; no solid ash remains, but charcoal always leaves a little ash, which proves it to be impure carbon.

OTHER IMPURE FORMS RESEMBLING CHARCOAL. — Charcoal is only one of the many common forms of carbon. Among those most nearly like charcoal we may mention now the *hard coal*, which is taken from mines for fuel; *coke*, the black and porous solid left in the retorts of gas-works; *bone-black*, obtained by heating bones in close vessels; *soot*, to be found in chimneys; and *lamp-black*, so much used in the manufacture of printing-ink.

**Lamp-black.** — Lamp-black is made by burning pitch

or tar with little air. The pitch is put into an iron pot and heated as shown at the left in the cut (Fig. 49). A dense black smoke is carried by the draught over into a large chamber. The blackness of this smoke is due to fine particles of carbon. This substance collects on the walls of the chamber, and is afterwards taken out as a fine black powder. This is lamp-black.

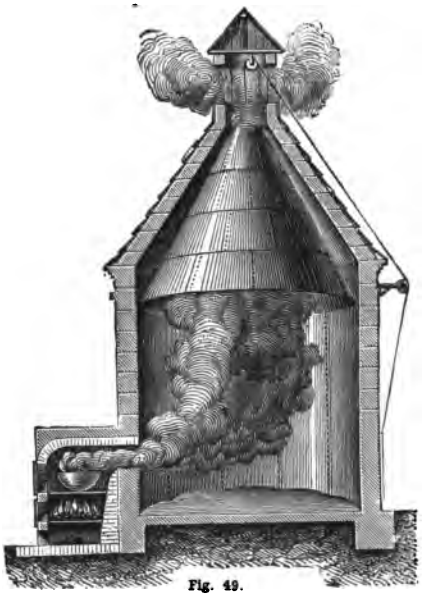


Fig. 49.

Printers' ink is a mixture of lamp-black and oil; every printed mark on this page is a thin layer of carbon clinging tightly to the paper.

**Action of Charcoal on Gases.** — Charcoal is very porous, and has remarkable power to absorb gases. Let us study this action by experiment with ammonia.

*Ex. 67.* — I must first heat the charcoal to redness to drive out the air already in it. I select a piece that has been well burned, make it about one inch long, and small enough to slide easily into my graduated cylinder. I place it in the bottom of my porcelain dish, or better, in a large iron spoon, cover it completely with fine sand, and put it in a good fire. After it has been heated to full redness for some time I let it cool, still protected by the sand.

While the spoon is cooling I fill my cylinder with ammonia gas, in this way:

Place the cylinder, inverted, through the ring of the support. Pour 5 cc. liquid ammonia into the side-neck tube, and put the delivery tube up into the cylinder. Then gently heat the liquid. I can hold it by means of a strip of paper which I wind around it, as shown in Fig. 50. Ammonia gas will be driven over into the cylinder, and, being little more than half as heavy as air, it will rise to the top and gradually expel the air until the cylinder is filled.

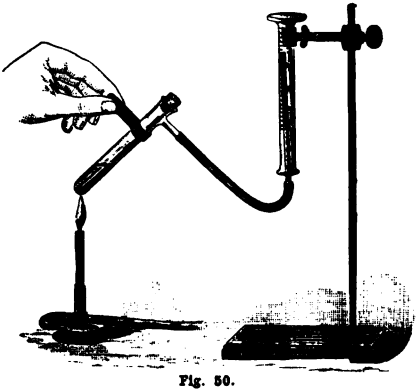


Fig. 50.

I now place the cold charcoal in the mouth of the cylinder and then quickly lower it into a dish of mercury. If the mercury rises in the cylinder, as in Fig. 51, it will prove that the ammonia gas is being absorbed by the charcoal.

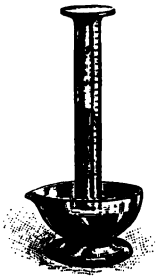


Fig. 51.

FACTS. — Charcoal will take up 90 times its own volume of ammonia gas, but only 8 or 9 times its volume of oxygen or nitrogen or carbon dioxide. This remarkable power of charcoal makes it very useful in hospitals and other places where offensive odors are to be found. It will absorb the bad gases, and thus purify the air.

Even animal substances, when decaying, lose their power to offend us by their odor, if covered with a layer of good charcoal. The decay will go on, but the odor will be lost.

**Action of Charcoal on Colors.** — Charcoal also has the power to absorb many coloring matters. Animal charcoal, or bone-black, possesses this property in higher degree than wood charcoal.

*Ex. 68.* — I prepare a filter, and rest the funnel in the mouth of the cylinder. I fill the filter nearly full of bone-black. Finally, I pour upon it some water colored with blue litmus. If it comes through still colored, I pour it back and let it run through a second time. Is the color removed? It has been absorbed by the charcoal.

*Ex. 69.* — In the same way I filter some water colored with cochineal.

*Ex. 70.* — I try a solution of dark brown sugar.

*Ex. 71.* — I try a solution of potassium chromate. This is a mineral coloring-matter; the others have been organic.

Is the color discharged in all these cases?

APPLICATION. — This power of charcoal makes it very useful for taking the color out of liquids. Great quantities are used to remove the brown color from crude sugar. After passing the dark syrup through a charcoal-filter it is bright and colorless.

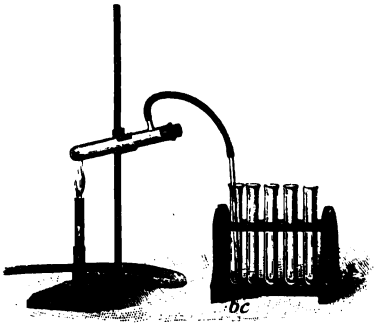


Fig. 52.

Charcoal filters are much used to purify drinking-water. The charcoal gradually loses its power as its pores are more and more filled with the impurities. Fresh portions should then be put in its place, or the same portion may be again heated to redness, which will restore its power.

**Action of Charcoal on Oxides.** — Charcoal has a strong attraction for oxygen, and when very hot it will decompose many compounds of oxygen in order to unite with it.

*Ex. 72.* — I make a mixture of 1 g. copper oxide with about its own bulk of powdered charcoal, put it into the side-neck ignition-tube, and place the end of the delivery tube in some lime-water contained in a test-tube, as shown in Fig. 52, and then apply the heat of the Bunsen lamp.

What is the effect on the lime-water?
What, then, was the gas produced by the action?
What are the elements of this gas?
In what form were these elements in the mixture?
Can you see any change in the black copper oxide?
What, then, must have been the action of the carbon?

THE FACTS. — The black copper oxide is composed of copper and oxygen, but when heated the carbon takes the oxygen away from the copper and unites with it to form carbon dioxide. This leaves the copper free. The carbon dioxide whitens lime-water. The copper left behind should appear distinctly reddish in the tube.

This power of carbon to decompose oxides makes it very useful in the work of getting metals out of ores. Iron, for example, is found in the mine in the form of iron oxide, and by mixing this ore with coal, and then heating it intensely in a furnace, its oxygen is taken away by the carbon, and the iron is left in the metallic form.

**The Diamond.** — The *diamond*, most brilliant of gems, is nothing but carbon. It is crystallized carbon. Dull, black charcoal, very common and very cheap, and the beautiful diamond, most costly of precious gems, are only two different forms of the same element.

The diamond is the hardest known substance. It cannot be cut or even scratched by any other. Very small

and otherwise useless diamonds are set in the end of a proper handle and are commonly used for cutting glass.

When found in the earth the diamond has the shape and appearance of a roughly rounded pebble. This rough gem is "cut" into one of two principal forms of jewel. They are called the *brilliant* and the *rose*. The first of these is regarded as the finest. Its shape is well shown in Fig. 53. Look at the gem sidewise, and it appears as seen in the upper part of the picture; look at the top of it, and it appears as seen in the lower part. The form of the *rose* is seen in Fig. 54; a side view above and a top view below.



Fig. 53.

**Graphite.** — Another form of carbon is known under the common name of *black-lead*, and is very familiar to us in its most useful shape, the lead of our lead-pencils. This name does not belong to it properly, for it is not lead. It has no properties like those of lead except that which allows it to leave marks upon paper. It is sometimes called *plumbago;* but the name by which it is generally known in chemistry is *graphite.*

Graphite is found in the earth. It is as black as coal, but it has a dull shining appearance, which coal has not. It is among the softest minerals to be found, and as per-

fectly opaque as can be. How unlike the hard and transparent diamond in these respects!

THREE FORMS OF THE SAME ELEMENT. — All three of these forms of carbon are alike in some respects. No fire can melt them. No liquid can dissolve them, if we except melted iron, which seems to dissolve a little carbon. They cannot be changed by exposure to the atmosphere; they suffer no decay, no rust. But let them be heated hot enough, where oxygen is present, and they will burn, and the experiment shows that they are not only much alike in some respects, but that they are actually the same kind of matter.



Fig. 54.

For we have proved by experiment, that when charcoal burns *carbon dioxide* is produced. By heating the diamond to a white heat, in oxygen gas, it will be consumed, and *carbon dioxide* will be found in its place. Graphite has also been burned, and the same substance, *carbon dioxide*, produced by its combustion.

Now we know that it takes carbon and oxygen always, and nothing else, to make carbon dioxide. The carbon must come from the fuel which burns. Hence charcoal, diamond, and graphite are all the same element, — carbon.

What other element have we found to exist in more than one form? What name is given to this property of elements?

How this element has come to be in such wonderfully

different forms we cannot tell. No chemist can change charcoal into diamond, nor can any one tell us how it has been done in the great laboratory of nature.

### CARBON DIOXIDE.

**Preparation.** — Carbon dioxide is the chief compound of carbon and oxygen. It is most easily obtained by the action of hydrochloric acid on marble.

*Ex. 73.* — My apparatus is much the same as that used in making oxygen (Ex. 23). It is shown in Fig. 55. A little lime-water is placed in the bottle *d*, and clear water in *a*, enough to cover the end of the glass tube. The conical flasks are joined by rubber tubes, the long tube in each to the short tube of the one before it, and then the long tube of *a* to the side-neck of the flask *s*. The joints are all air-tight.

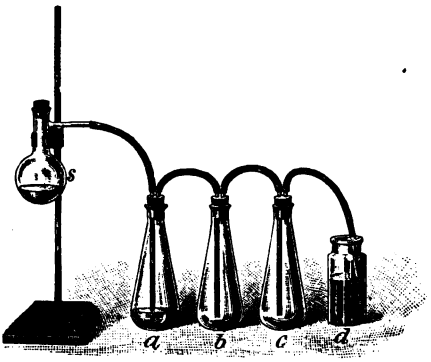


Fig. 55.

I put about 10 g. of marble in small pieces into the side-neck flask and pour upon it about 25 cc. of dilute hydrochloric acid, half water, and at once close the flask with its stopper. The gas, making in *s*, will drive the air before it all out at *d*. The flame of a burning splinter thrust into *d* will tell when the bottle is full of the carbon dioxide. Then exchange this bottle for another having some blue litmus-water instead of lime-water.

If the gas begins to come too slowly, I may remove the cork of *s*, add more acid, and replace it. Now consider the following things:

Describe the action which took place in *s*.
What effect was produced in the lime-water?
What is the effect of this gas on flame?
What is the effect of it on blue litmus?
Why does the gas stay in the open bottle?

*Ex. 74.* — What if the vessel of carbon dioxide is open downward? I remove the tubes and cork from flask *c*, Fig. 55, and then hold the flask's mouth upon the lip of a small wide-mouth bottle, as if to pour the gas, as shown in Fig. 56. After a minute I place the flask upon the table and plunge the flame of a splinter down into the bottle, and afterward into the flask. The flame shows that the gas was actually poured, like water, from *c* into the bottle.



Fig. 56.

*Ex. 75.* — Is this gas soluble in water? The gas bubbled through the water in *a*, Fig. 55, and I test that water to see whether there is carbon dioxide in it. I put some lime-water in a test-tube, and, after taking the rubber tubes from *a*, I lift the flask and pour water from it into the lime-water, as shown in Fig. 57.

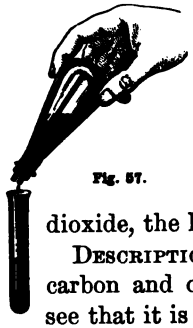


Fig. 57.

If the water in *a* has dissolved carbon dioxide, the lime-water will be whitened.

DESCRIPTION. — Carbon dioxide is a compound of carbon and oxygen: this was proved by Ex. 26. We see that it is a colorless gas, and much heavier than air (Exs. 73, 74). In fact it is about one and one-half times heavier than air (1.529). It whitens lime-water by combining with the lime and making the white solid called *calcium carbonate*, which the water cannot dissolve, and it

reddens blue litmus-water (Ex. 73), for a reason which will appear by and by.

This gas will put out a fire (Ex. 73) just as quickly as will water, and for the same reason. A body will not burn without oxygen, and when covered with water the oxygen of the air cannot get to it. So when in carbon dioxide, no oxygen can reach the body, and the fire dies.

So, too, this gas will cause death, just as will water. An animal must have oxygen to breathe, or it must die. It drowns in water because water keeps the air away from its lungs; carbon dioxide will do the same thing. There is too little of this gas in the open air to do harm, but where it collects in large quantities an animal dies because it can get too little oxygen to keep it alive.

It sometimes collects in mines, and the miners call it *choke-damp*. It sometimes collects in wells, and makes them dangerous to enter; a lighted candle will tell a workman whether this gas is present in dangerous quantities. It often collects in school-rooms and churches and other unventilated houses, but not often enough of it to do much harm. The truth is that the breath of people is full of other gases, some of which are very poisonous, and the mischief in unventilated rooms is done by these companions of the carbon dioxide instead of by this gas itself.

Carbon dioxide is soluble in water (Ex. 75). It is found that water will usually dissolve about its own volume. But, like all gases, more will dissolve in colder water, and still more when under greater pressure. Nearly all water contains this gas, but the quantity in solution is generally small. In some springs, however, it is very abundant. It is so in the noted mineral springs of Saratoga. And the refreshing summer drink — soda-water — is nothing but water charged with a large quantity of this gas which has been forced into it by great pressure.

**Carbon Monoxide.** — Besides carbon dioxide there is another compound of carbon and oxygen called *carbon monoxide.* See p. 82. It is a colorless gas. It takes fire easily, and burns with a pale blue flame; this is the flame which is often seen playing over the surface of a new-made coal-fire. When the coal burns at the bottom of the grate it produces carbon dioxide, as usual, but as this carbon dioxide goes up through the hot coal above it gives up a part of its oxygen and becomes carbon monoxide. And then when this monoxide comes out into hot air, at the surface, it takes oxygen and becomes carbon dioxide again.

This gas is very likely to escape being burned. There may not be air enough, or there may not be heat enough to burn it. This should be remembered when coal is used for warming rooms, for carbon monoxide is a very poisonous gas. Accidents have many times happened from burning charcoal in open fires with poor draught. This poisonous gas escaping into the room destroys life.

COMPOSITION OF THE TWO OXIDES. — It is found by analysis of the two gases that the carbon monoxide contains just one-half as much oxygen as the carbon dioxide. The proportions are as follows:

In carbon monoxide there are 16 of O combined with 12 of C
" " dioxide " " 32 " O " " 12 " C

What law does this illustrate?

**Compounds of Carbon and Hydrogen.** — No other elements form so many compounds with one another as do carbon and hydrogen. Of carbon and oxygen there are only two. Of nitrogen and oxygen there are five, and this is rather an unusual number for two elements only. But of carbon and hydrogen the compounds are many scores in number. They are called *hydrocarbons.* Just now we will study only one of these.

**Methane.** — This is the simplest one of the hydrocarbons. It is commonly called *marsh-gas*, which is a very appropriate name, because it is abundant in marshy places. Have you never seen the water of ponds and quiet pools stirred by bubbles of gas breaking at the surface? It is a very common thing to be seen in water standing over muddy bottoms. The bubbles are of marsh-gas. They are set free by the decay of vegetable substances in the mud below.

Marsh-gas flows from coal-beds into the mines where the laborers are at work. The most awful effects are then sometimes produced by it. It is very combustible, and, more than this, when mixed with air it is terribly explosive. Fancy it flowing into a mine, mixing with the air around the miners, and the mixture then touching the flame of the miner's lamp! The lamps and the lives of the miners are then extinguished at once by a terrible explosion.

The miners have given this gas the name of *fire-damp*.

It is found by analysis that methane is made of carbon and hydrogen only, and in quantities by weight which are as 12 to 4.

In methane there are 12 parts of C combined with 4 parts of H.

How many combining weights of carbon in the 12 of C?

How many combining weights of hydrogen in the 4 of H?

## ELEMENTS, MOLECULES, AND ATOMS.

**The Number of the Elements.** — We have already seen, p. 24, that an element is a substance which has never yet been decomposed.

At present (July, 1886) seventy-one elements are known. Some of these have been discovered very lately, and perhaps others may be found in the future. On the other hand some so-called elements have been found to be compounds, and it may be that some of the seventy-one shall yet be decomposed. All that we can say is that there are seventy-one kinds of matter which chemists to-day are not able, by any means in their power, to break into simpler bodies. As out of twenty-six letters of the alphabet all the words in the English language are made, so out of these seventy-one elements all the compounds in nature are formed.

But only a few of this small number are at all abundant. In fact, the larger part of the earth, and all it contains, is made up of only about a dozen elements. All the water on the globe consists of oxygen and hydrogen. Four-fifths of the air is nitrogen. These three, with carbon, make up by far the larger part of all plants, and the bodies of animals. And as for rocks and soils, they consist chiefly of those just named, with eight others, — sulphur, silicon, potassium, sodium, calcium, magnesium, aluminum, and iron. With these exceptions, the elements are not common, and one-third of the whole number are very rare.

The names of all the elements now known are given in the following table. The symbols and atomic weights will be explained soon.

## THE SEVENTY-ONE ELEMENTS.

| NAMES. | Symbols. | Atomic Weights. | NAMES. | Symbols. | Atomic Weights. |
|---|---|---|---|---|---|
| Aluminum . . | Al. | 27.3 | Molybdenum . | Mo. | 96. |
| Antimony . . | Sb. | 120. | Nickel . . . | Ni. | 58. |
| Arsenic . . . | As. | 75. | Nitrogen . . . | N. | 14. |
| Barium . . . | Ba. | 137. | Osmium . . . | Os. | 198.6 |
| Beryllium [1] . | Be. | 9. | Oxygen . . . | O. | 16. |
| Bismuth . . . | Bi. | 208. | Palladium . . | Pd. | 106. |
| Boron . . . | B. | 11. | Phosphorus . | P. | 31. |
| Bromine . . . | Br. | 80. | Platinum . . | Pt. | 195. |
| Cadmium . . | Cd. | 112. | Potassium . . | K. | 39.1 |
| Cæsium . . . | Cs. | 133. | Rhodium . . | Rh. | 104. |
| Calcium . . . | Ca. | 40. | Rubidium . . | Rb. | 85.5 |
| Carbon . . . | C. | 12. | Ruthenium . . | Ru. | 103.5 |
| Cerium . . . | Ce. | 141. | Samarium . . | Sm. | 150. |
| Chlorine . . . | Cl. | 35.5 | Scandium . . | Sc. | 44. |
| Chromium . . | Cr. | 52. | Selenium . . | Se. | 79. |
| Cobalt . . . | Co. | 59. | Silicon . . . | Si. | 28. |
| Columbium [2] . | Cb. | 94. | Silver . . . . | Ag. | 108. |
| Copper . . . | Cu. | 63.3 | Sodium . . . | Na. | 23. |
| Decipium . . | Dp. | ? | Strontium . . | Sr. | 87.5 |
| Didymium [3] . | Di. | 142.3 | Sulphur . . . | S. | 32. |
| Erbium . . . | Er. | 166. | Tantalum . . | Ta. | 182. |
| Fluorine . . . | F. | 19. | Tellurium . . | Te. | 125. |
| Gallium . . . | Ga. | 69. | Terbium . . . | Tb. | ? |
| Germanium. . | Gr. | 72.75? | Thallium . . | Tl. | 204. |
| Gold . . . . | Au. | 196.5 | Thorium . . . | Th. | 232. |
| Hydrogen . . | H. | 1. | Thulium . . . | Tm. | ? |
| Indium . . . | In. | 113.6 | Tin . . . . | Sn. | 118. |
| Iodine . . . | I. | 127. | Titanium . . | Ti. | 48. |
| Iridium . . . | Ir. | 193. | Tungsten [4] . | W. | 184. |
| Iron . . . . | Fe. | 56. | Uranium . . . | U. | 240. |
| Lanthanum . | La. | 138.2 | Vanadium . . | V. | 51.2 |
| Lead . . . . | Pb. | 207. | Ytterbium . . | Yb. | 173. |
| Lithium . . . | Li. | 7. | Yttrium . . . | Yt. | 89. |
| Magnesium . . | Mg. | 24. | Zinc . . . . | Zn. | 65. |
| Manganese . . | Mn. | 55. | Zirconium . . | Zr. | 90. |
| Mercury . . . | Hg. | 200. | | | |

[1] Beryllium is also called Glucinum, with the symbol Gl.

[2] Columbium is also called Niobium, with the symbol Nb.

[3] The announcement has come from Vienna that Didymium has been decomposed. See *Chemical News*, vol. 52, and *Nature*, vol. 32, p. 435.

[4] Tungsten has also been called Wolframium.

**The Three Forms of Matter.** — Earth, water, and air represent the solid, liquid, and gaseous forms of matter. So far as we know, all kinds of matter are in one or another of these three conditions. But ice and water and steam are all one kind of matter, and yet one is a solid, another is a liquid, and another a gas.

Ice, when heated, melts into water, and water when heated to a higher temperature boils into steam. We find that many other solids are like ice in this respect — they melt when heated, and that other liquids are like water — they may be changed into vapor by heat.

These are *facts*. We know them to be so, because we have seen these changes happen over and over again. Indeed, so many substances have been tried and found to be solid at low temperature, liquid at some higher temperature, and gaseous at some temperature higher still, that we feel quite sure that the facts are the same for all; and so we say that

*The solid, liquid, and gaseous forms of matter depend on the amount of heat in them.*

This one statement, made out of many facts, is a *law*.

**Another Effect of Heat.** — If we warm a piece of iron it will become larger, and if we cool it it will become smaller. The same thing is true of a piece of wood; it is larger when it is hot than when it is cold. A given weight of water, or of air, is found to be larger and larger as it is heated more and more. These are *facts*. They have been proved to be true of the substances named, and of a great many more beside. Now let us put all such facts together into one statement by saying that

*Heat makes a body larger without changing its weight.*

This is a *law*. A law, in science, is simply a single statement which covers a large number of facts.

THE REASON. — A body is larger when it is hot; can we tell why it is so? Nobody knows with absolute certainty why bodies should expand when heated. But the best way to account for it is to suppose that matter is made up of very small particles which are quite separate from one another, so that they may be driven farther apart or crowded nearer together. Heat drives them farther apart and makes them fill more space, and cold allows them to fall more closely together, and thus to fill a smaller space. Surely this is a very good reason why heat should make a body larger, and cold should make it smaller. If a body were made up of separate particles it would behave just as we see that it does behave when heated and cooled, and so we may suppose that it is made up of such particles. *Perhaps* it is so.

If we imagine every body to be made up of small particles which are quite separate, we can understand many other facts beside expansion by heat. Take, for example, the fact that you make an india-rubber ball smaller when you squeeze it. If we are right in supposing that the ball is made up of a great multitude of small, separate particles, the hand would make it smaller by crowding them closer together. The ball acts *just as if it were* made up of separate particles. *Probably* it is so.

In a great many other ways bodies behave exactly as if they were made up of such separate particles. Indeed, so far as they have been examined, the facts are *always* just what we should expect if matter be made up of small particles which are some distance apart. On this account we *believe* it to be so.

Now, something which is believed because it gives good reasons for a large number of facts is called a *theory.* That all bodies are made up of minute particles which do not touch one another is our theory of matter.

The student cannot too early nor too carefully fix in his mind the difference between *facts* and *theories*. Facts are things *known* to be true; theories are things *believed* to be true.

**Molecules.** — According to the *theory* of matter just stated, all bodies are made up of separate particles. These particles cannot be seen. They are so small as to be quite beyond our sight with the best microscope. They are the very smallest pieces of a substance which can possibly exist, so that to divide them into parts will change the substance into something else. They are called *molecules*. A molecule is a particle of substance so small that it cannot be divided without changing the nature of the substance. Molecules of the same substance are all alike, but the molecules of different kinds of matter are not alike.

**Some Facts about the Expansion of Gases.** — It has been found by very careful and repeated experiments that if we take 273 cc. of air as cold as freezing water, 0° C. and warm it just 1° C., it will become just 274 cc. Add another degree of heat and it will become 275 cc., and so on, each degree of heat expanding it just 1 cc. In other words, one degree of heat will always expand air just $\frac{1}{273}$ of what its volume is at 0° C.

The rate has been found also for many other gases, and *it is the same for all.* This is a very important *fact.* It is not so among solids and liquids; no two solids expand exactly alike; no two liquids. Why do all gases expand alike when heated, while solids and liquids do not?

Again, if air is pressed in a close vessel it will be crowded into smaller volume; but, on the other hand, if we put less pressure on it the air will expand. Other gases behave in the same way. And the same change of pressure will change the volume of one just exactly as much as another. This is another important *fact.* It is

not a fact for solids and liquids? Why should the same pressure condense all gases exactly alike; and the removal of the same pressure expand all gases exactly alike?

**The Theory.** — These and other facts led Avogadro, in 1811, to suppose that when two gases are equally warm, and under the same pressure, a cubic inch of one *contains just as many molecules* as a cubic inch of the other. And since all gases behave, in so many respects, just as they would be expected to do if this famous hypothesis were true, it is believed to be as Avogadro supposed.

*Equal volumes of all gases, at the same temperature and pressure, contain the same number of molecules.*

This is Avogadro's *theory:* but it is often called Avogadro's law.

**Chemical Changes are Changes in the Molecules.** — When water is decomposed by electricity (Ex. 37) it is broken into two parts; one part is hydrogen, the other is oxygen. The change is believed to take place in each separate molecule. When mercuric oxide is heated (Ex. 5) every molecule of it is broken into two parts; one is mercury, the other oxygen. When marble and hydrochloric acid act (Ex. 73) every molecule of these two substances is broken into two parts, and then those parts fall together again, in a different way, and make the molecules of two new kinds of matter.

**Atoms.** — Now what shall we call the smaller particles in a molecule? They are *atoms.* These are the very smallest pieces of matter that take part in a chemical change. They cannot be divided by any known process whatever. All bodies are made up of molecules, and molecules are made up of atoms.

If the atoms of a molecule are alike, the substance is an *element,* but if the atoms in a molecule are not all alike the substance is a *compound.*

**The Explanation of the Law of Multiple Proportions.** — The law of multiple proportions is the statement of a *fact.* We know that the combining weight of an element is never divided in chemical changes; can we tell why?

The most natural reason is found by supposing the combining weight to be the weight of a particle of matter which cannot be divided, that is, of an *atom.* Such particles will pass unbroken from one kind of compound into another. One of them may go alone, or some whole number of them may go together. The weight of one of them is the combining weight of the element, and of any whole number of them, is a multiple of that combining weight. And so —

If one element combines with another, in more proportions than one, these proportions must be multiples of its combining weight.

**The Atomic Theory.** — This theory of matter is known as the *Atomic Theory.* Its main points are these: All bodies are made up of molecules, and molecules are made up of atoms. All chemical changes are changes in the arrangements of the atoms, and the combining weights of the elements are the relative weights of their atoms.

The theory of atoms was first proposed in 1808 by John Dalton. He was the first to see clearly that the combination of the elements, in definite and in multiple proportions, are facts which seem to show that there are particles of matter which cannot be divided. The atomic theory was made to explain these laws of combination, and it is well fixed in chemistry, because it not only explains these, but has been found to explain new facts, as fast as they have been discovered, down to the present time.

**Symbols of Elements.** — Instead of writing the full name of the element hydrogen, we may simply write the first letter, H, to represent it. H is then called the

*symbol* of hydrogen. In the same way O is the symbol of oxygen, N of nitrogen, and C of carbon.

Sometimes the names of two or more elements begin with the same letter, as carbon, calcium, and copper. In this case a small letter is used with the large initial, thus: the symbol of calcium is Ca, and that of copper is Cu. The latter is from the Latin name of copper, which is Cuprum. *A symbol represents the name, and also just one atom of an element.*

**Formulas of Compounds.** — Instead of writing the full name mercuric oxide, we may simply write the symbols of the elements of this compound, side by side, Hg for mercury and O for oxygen, making Hg O to represent it. Hg O is then called the *formula* for mercuric oxide. H Cl is the formula for hydrochloric acid.

*The formula of a compound represents one molecule of it.*

$H_2O$ is the formula for water; the small figure 2 means that there are two atoms of hydrogen combined with one of oxygen in the molecule. $CO_2$ is the formula for carbon dioxide, and shows that one molecule contains one atom of carbon and two atoms of oxygen. $H_2SO_4$ is the formula for sulphuric acid; what does it show? $3\,H_2SO_4$ means three molecules of sulphuric acid, and $2\,H_2O$ means two molecules of water. Notice the different meaning of figures placed before the formulas and those placed at the right and a little below the symbols, in the formula.

In reading symbols and formulas, the names of the substances, and not the letters used, should always be given. We should use symbols and formulas to shorten writing, but not to shorten speech.

**Atomic Weights.** — We have learned that hydrochloric acid contains 35.5 times as much chlorine as hydrogen; we may say this of any quantity; it is true of one molecule, which contains one atom of chlorine and one of hydrogen.

These numbers, 1 and 35.5, are called the *atomic weights* of hydrogen and chlorine, because they are supposed to represent the relative weights of the atoms of these two elements. We have no idea how much, or rather how little, the atom of hydrogen really weighs, but we do believe that, whatever it does weigh, the atom of chlorine weighs 35.5 times as much. We cannot tell the weight of either, in fractions of a gram or ounce, but we can call the weight of the hydrogen atom 1, without saying what, and then we can say that an atom of chlorine weighs 35.5.

These numbers are the same as those which have been called combining weights, p. 98. We *know* that they are the smallest relative weights of these elements which combine; this is a *fact* discovered by experiment. It is *believed* that they represent the weights of the atoms of these elements. This is a part of the atomic *theory.*

And now we see that the symbol of an element not only represents one atom, but also one atomic weight.

**Molecular Weights.** — If a molecule is made of atoms we can find its weight by adding the weights of its atoms together. There are three atoms in a molecule of water, $H_2O$, two of hydrogen and one of oxygen. The weight of an atom of hydrogen is 1, and of an atom of oxygen is 16, and the weight of all three atoms together must be 18. The molecular weight of water, $H_2O$, is therefore 18. It is not 18 grams, nor 18 grains, but simply 18 *times as much as the weight of an atom of hydrogen.*

The atomic weights are given in the table on p. 118, and these will enable one to find the molecular weight of any substance whose formula is known. Try it in the following examples.

What is the molecular weight of hydrochloric acid, $HCl$?
" " " " sodium chloride, $NaCl$?
" " " " mercuric oxide, $HgO$?
" " " " sulphuric acid, $H_2SO_4$?

***Reactions.*** — Chemical changes are also called *reactions*. Now chemical changes are changes in the molecules of substances, and we can show what they are by making changes in the formulas which represent the molecules. In this way symbols and formulas are a great help in the study of experiments.

For example, suppose we wish to know all about the reaction of sodium and hydrochloric acid. We must begin by making the experiment. We may then study the results by the help of symbols and formulas.

REACTION OF SODIUM AND HYDROCHLORIC ACID. *Ex. 76.* — I measure 2 cc. of strong hydrochloric acid into a test-tube which stands in the rack, drop in upon it a piece of sodium as large as a small pea, and loosely cover the mouth of the tube with a piece of paper. After half a minute, I push the flame of a lighted match down into the tube. When the sodium has disappeared, I drop in another piece of about the same size.

What gas is set free by the chemical action?

What proof is seen that another substance is produced?

To find out what is this second product, I first let it settle to the bottom of the tube, and then pour off the liquid so carefully that I leave the solid behind almost dry. In this way I get rid of most of the acid which was not decomposed. I then add three or four cubic centimeters of water, to dissolve the white solid, pour the solution into a small porcelain dish, and heat it over a small flame until the water is all driven off and the solid remains dry and white. The dish must now stand until cold, after which I add a few drops of water in which the white solid again dissolves. By tasting this liquid you may learn

What is this second product of the chemical action.

We have now got the facts about the reaction we are studying. To get the facts is always the first step in an investigation.

NEXT WRITE THE REACTION. — Now that we know what substances were used and what new substances were made, the next step is to use formulas and signs to show what happened.

We added sodium to the hydrochloric acid in the tube; we can show that fact by writing the symbol of sodium, which is Na, and the formula of the acid, which is H Cl, with the sign of addition between them, thus: Na + H Cl. We know by the taste, and by the explosion, that common salt and hydrogen were produced. We can show this by writing the formula of common salt, or sodium chloride, Na Cl, and the symbol of hydrogen, H, with the sign of addition between them, thus: Na Cl + H. We know that there is no loss nor gain of matter in any chemical change. The new substances must contain every atom of the old, and no others. We can show this by writing the old substances and the new, with the sign of equality between them. Thus we have —

$$\mathrm{Na + H\,Cl = Na\,Cl + H}$$

We read this reaction in this way: one atom of sodium, with one molecule of hydrochloric acid, yield one molecule of sodium chloride, and one atom of hydrogen. This describes the chemical change which took place.

NOW USE ATOMIC WEIGHTS. — This chemical expression *may be changed into numbers* by writing the atomic weight of each element, under its symbol, wherever that symbol appears. Thus —

$$\begin{array}{ccccccc} \mathrm{Na} & + & \mathrm{H\,Cl} & = & \mathrm{Na\,Cl} & + & \mathrm{H} \\ 23 & + & (1+35.5) & = & (23+35.5) & + & 1 \\ 23 & + & 36.5 & = & 58.5 & + & 1 \end{array}$$

If we had to make a quantity of common salt by this process this last equation would be very useful because it tells us just how much of the substances to use. We may

call these numbers grains or grams or pounds, or any other weights we please. The equation shows that if we use 23 grams of sodium and 36.5 grams of hydrochloric acid, we can get just 58.5 grams of salt and 1 gram of hydrogen. Knowing this it is easy to find how much of the constituents to take for any desired quantity whatever of the compound.

FINALLY SOLVE A PROBLEM. — How much sodium would be needed to make 100 g. of common salt by reaction with hydrochloric acid?

If for 58.5 g. of salt we need 23 g. of sodium, then for 100 g. of salt we should need $\frac{100}{58.5} \times 23$, which is 39.31+.

ANOTHER EXAMPLE. — Let us study the decomposition of mercuric oxide by heat. We have made the experiment already. In Ex. 5 it was found that mercury and oxygen were the products. We take the symbols from the table, p. 118, and write the reaction thus:

$$Hg\,O = Hg + O$$

Read this equation in full. Refer to the table of elements, p. 118, for the atomic weights, and put them under the symbols. What does this new equation show?

How much mercury could be obtained from 100 g. of the oxide? And how much oxygen?

## ACIDS, BASES, AND SALTS.

**Acids.**—We have already made use of several substances called acids. We now set out to discover just what is meant by the term acid. Let us examine a few of these substances for this purpose.

*Ex. 77.* — I fill a bottle two-thirds full of water and add enough solution of blue litmus to color it distinctly. I next add dilute sulphuric acid, 1 of acid to 10 of water, drop by drop, and note the change in color.

I now add a drop or two of the same acid to a test-tube nearly full of water, and touch a drop of this dilute solution to my tongue, and note the taste of the acid.

I finally pour 5 cc. of strong sulphuric acid into 40 cc. of water in a bottle, drop into it a clipping of sheet-zinc, and cover the bottle with a square of heavy paper. In a few moments I bring a lighted match to the mouth of the bottle, lifting the cover at the same time, and note the combustion which takes place. What is this gas which is set free by the metal?

*Ex. 78.* — I now use hydrochloric acid instead of sulphuric, and again note the

Change in the color of the blue litmus.
Taste of the acid.
Action with zinc.

*Ex. 79.* — I next try acetic acid, and use magnesium instead of zinc. I put 10 cc. of the acid in a test-tube, and drop upon it a short piece of magnesium ribbon. Note whether the same gas is given off as in the other cases. Note also the taste of the acid, and its action on blue litmus.

*Ex. 80.* — I repeat the experiment with acetic acid, but use iron instead of magnesium. I add the iron in the form of filings, or the smallest tacks, and *gently* heat the liquid. Is there any gas set free? Is it hydrogen?

*Ex. 81.* — I use acetic acid again, but this time I add small clippings of zinc, and heat the acid. Is the action the same as before?

We find that sulphuric acid, hydrochloric acid, and acetic acid are alike in three things: they are all sour to the taste, will redden blue litmus, and yield hydrogen by the action of a metal. There are many other substances having these same characters. All such are called *acids*.

THE CHIEF CHARACTERISTIC. — The following table shows the composition of several acids, by formulas:

| | |
|---|---|
| Nitric acid | $HNO_3$ |
| Sulphuric acid | $H_2SO_4$ |
| Phosphoric acid | $H_3PO_4$ |
| Hydrochloric acid | HCl |
| Hydroiodic acid | HI |
| Hydrobromic acid | HBr |

and a glance at these symbols shows that hydrogen, H, is a constituent in every one. *Every acid contains hydrogen.*

But there are many compounds containing hydrogen which are not acids. An acid is a compound containing hydrogen, *which may be driven out by a metal*, as in our experiments. This is the chief character of an acid.

THE TWO CLASSES. — Look at the two sets of formulas in the table. In each acid of the first set, oxygen, O, is present; it is not to be found in those of the second.

Now we may learn from this that there are two classes of acids. One class contains the element oxygen, the other does not.

**Salts.** — When in Ex. 77 our zinc drove the hydrogen out of sulphuric acid, zinc sulphate was also made.

If we write the reaction —

$$Zn + H_2SO_4 = H_2 + ZnSO_4$$

it shows how zinc and sulphuric acid may yield hydrogen and zinc sulphate.

We see that 1 atom of the metal, Zn, *takes the place of* 2 atoms of hydrogen, $H_2$, in the molecule of acid, so that instead of $H_2SO_4$ we have $ZnSO_4$.

We should be careful to notice that the metal puts itself in the place of the hydrogen of the acid to form the new molecule.

The action is the same when zinc acts on other acids instead of sulphuric. The zinc drives hydrogen out and puts itself in the place of it.

Many other metals have the same power to act upon acids. Not every metal can do this to every acid (Ex. 81), but the power to put itself in the place of hydrogen in acids is very common among the metals.

The new compounds, formed by this action of the metals, are called *salts*. When sodium and hydrochloric acid are used, the new compound, as we have seen, is common salt; you have only to evaporate the liquid to get the familiar white solid. Every *acid* would yield a different kind of salt with sodium. Each different *metal* also yields a different salt with every acid on which it acts.

DEFINITION. — A salt is *a compound formed by putting a metal in the place of hydrogen in an acid.*

**Hydroxides.** — Some metals act on water very much as the metals act on acids, but the new compounds will be of a quite different character. Thus:

*Ex. 82.* — I fill a bottle half full of water; add just enough litmus to color it distinctly; change its color to red

by stirring it with a rod moistened with an acid, and then drop upon its surface a bit of sodium. Note the change in color, and prove that hydrogen is set free.

*Ex. 83.* — I repeat the experiment, but use 25 cc. of water without the litmus. When the sodium is gone, I examine the liquid in three ways:

1. I moisten the end of a rod with the liquid, touch my tongue with it, and note the taste.

2. I try its action on blue litmus-water. I then try it on *red* litmus-water in this way: I take a little water in a tube, blue it with litmus, and then redden it with a rod moistened with acid. To this I add a few drops of the liquid, and note the change in color.

3. I put the rest of the liquid in a porcelain dish and evaporate it to dryness, using a small flame toward the end of the operation, and note the appearance of the solid left behind.

THE FACTS. — The sodium decomposes the water, and the two new substances are hydrogen gas, and a white solid which is very "alkaline" to the taste, and able to restore the blue color to reddened litmus.

We may write the reaction in the usual way:

| Na | + | $H_2O$ | = | NaHO | + | H |
|---|---|---|---|---|---|---|
| Sodium | | Water | | Sodium hydroxide | | Hydrogen |

and this shows that 1 atom of the metal, Na, takes the place of 1 atom of the hydrogen, H, in the molecule of water, and makes a new molecule of what is called sodium hydroxide, NaHO. This sodium hydroxide is the white substance left behind in the dish.

It is named for the elements in it — sodium, hydrogen, and oxygen. The last two names are combined in hydroxide, and so we get the name sodium hydroxide.

In like manner we have potassium hydroxide, KHO, and ammonium hydroxide, AmHO.

*Ex. 84.* — I half fill two bottles with litmus-water reddened by acid, and then to one I add drops of potassium hydroxide (caustic potash) solution to the other drops of ammonium hydroxide (ammonia).

Note the change in the color of each.

Carefully learn the taste of each.

THE FACTS. — These three hydroxides are alike, very alkaline in taste, and alike able to restore the blue of reddened litmus. They restore the blue color by decomposing the acid which reddened it. In this way they *neutralize* the acids. This is their chief characteristic.

It has been found by experiment, that nearly all metals have compounds of this kind. The hydroxides are a very large class indeed.

DEFINITION. — A hydroxide is a compound of a metal with hydrogen and oxygen, which will neutralize an acid.

OTHER NAMES. — These substances are also called *hydrates;* instead of sodium hydroxide we may say sodium hydrate. They are also called *bases.*

**Reaction of Bases and Acids.** — But *how* does a base or hydroxide neutralize an acid? Take the sodium hydroxide and hydrochloric acid for example.

*Ex. 85.* — I put 15 cc. of hydrochloric acid into a bottle and drop into it a small bit of litmus-paper,[1] which instantly becomes red. I next add a solution of sodium hydroxide little by little, shaking or stirring the liquid well after each addition. I watch the color of the litmus-paper: it will after a while show signs of turning blue. I then add the hydroxide, a drop at a time, until after the last drop the paper remains blue. By this sign I know that the acid is neutralized.

[1] Litmus-paper is made by soaking filter-paper in a strong solution of litmus in water and then drying it; the paper has a deep blue color.

I next search for the product of the action. I put the liquid into a porcelain dish and evaporate it down carefully until dry. Let it cool, and then examine the residue. Note its color and its taste. By its taste we know what this new substance is.

We find that common salt (Na Cl) is one product of the action of these two liquids. By writing the reaction, we may discover that there should be another. For we must have new substances enough to contain all the atoms in the old ones. Thus, if we write —

$$\underset{\text{acid}}{HCl} + \underset{\text{base}}{NaHO} = \underset{\text{salt}}{NaCl} + \underset{\text{water}}{HHO \text{ or } H_2O}$$

we see that water must be a product of the action. In this case the acid and the base neutralize each other by forming a salt and water.

In other cases the action is the same. It is a fact that when an acid and a base act upon each other the result is to produce *a salt* and *water*. There are few exceptions (p. 95) to these important facts, viz.:

When a metal acts on an acid a *salt* and *hydrogen* are produced. When a base acts on an acid a *salt* and *water* are produced.

**Neutral Compounds.** — Water is a good example of bodies which are neither acids nor bases nor salts. It will not change the color of litmus either to red or blue. It is not sour to the taste like the acids, nor caustic like the bases, nor is it made by putting a metal in the place of hydrogen in the acids, as the salts are. There are a great many other compounds like water in these respects. Bodies which will not change the color of either blue or red litmus are called *neutral bodies*.

### CHEMICAL NAMES.

The names of chemical compounds are not invented to simply suit the fancy of the chemist; they are made by certain rules, and are meant to *show the composition of the substances to which they are given.*

**The Names of Acids.** — Sulphuric acid is made up of sulphur, oxygen, and hydrogen; its name is so made that the presence of these three things is shown.

The presence of sulphur is shown by using the name of that element. The presence of oxygen is shown by the ending *ic*. The presence of hydrogen is shown by the word *acid*, because all acids are known to contain this element.

Sulphurous acid is another compound of these same elements, sulphur, oxygen, and hydrogen. Sulphur is shown by its own name as before. Oxygen is shown by the ending *ous*. Hydrogen is shown by the word acid as before.

It will be seen that sulphuric acid and sulphurous acid are made of the same elements. The first contains a larger proportion of oxygen than the other, and this is their only difference in composition. Now *ic* is used in the name of the one that contains *the larger proportion* of oxygen, and *ous* in the name of the other.

In the same way we have two acids of phosphorus: they are called phosphor*ic* acid and phosphor*us* acid. The first contains a larger proportion of oxygen than the last. This fact is shown by using the ending *ic* in the name of the first, and *ous* in the name of the last.

In some cases more than two acids are made of the same elements. Prefixes are then used in addition to the endings *ic* and *ous*. These prefixes are *hypo* and *per*. There is, for example, the hyposulphurous acid. It contains a

less proportion of oxygen than the sulphurous acid, and this is shown by the prefix *hypo.* The prefix *per* means a larger proportion of oxygen. There is, for example, the perchloric acid. It contains a larger proportion of oxygen than the chloric acid: this is shown by the prefix *per.*

**The Names of Salts.** — The name of a salt is intended to show the names, of the metal and the acid, by which the salt is supposed to be made. If sodium act on sulphuric acid, hydrogen is liberated and a salt is at the same time formed; this salt is called sodium sulphate. It is easy to see that this name suggests the name of the acid and also of the metal. When sulphuric acid is used the salt which is made is called a *sulphate,* no matter what metal is employed.

But sulphurous acid yields a class of salts which are called sulphites. Notice the difference:

Sulphuric acid yields sulphates.
Sulphurous acid yields sulphites.

The same difference is found in other cases. A salt formed by chloric acid is called a *chlorate;* but one formed by chlorous acid is called a *chlorite.* Nitric acid yields *nitrates,* but nitrous acid yields *nitrites.*

The rule is this: if the name of the acid ends in *ic,* the names of its salts shall end in *ate;* but if the name of the acid ends in *ous,* the names of its salts shall end in *ite.*

**The Names of Bases.** — All bases are made up of hydrogen, oxygen, and a metal, and their names show this fact. The word hydroxide suggests hydrogen and oxygen, and hydroxide is the name of this whole family of compounds. Each individual member of this family is distinguished by the name of the metal which is in it.

The following examples will make this method clear. See how the names of the elements suggest the name of the base which they form, and how the name of the base also suggests the names of the elements of which it is made.

| | | | | | |
|---|---|---|---|---|---|
| Sodium, | hydrogen, | oxygen, | form | the | Sodium hydroxide. |
| Potassium, | " | " | " | " | Potassium hydroxide. |
| Calcium, | " | " | " | " | Calcium hydroxide. |
| Iron, | " | " | " | " | Ferric hydroxide. |
| Iron, | " | " | " | " | Ferrous hydroxide. |

There are in many cases two hydroxides of the same metal, and the endings *ic* and *ous* are used to distinguish them. The names of the iron hydroxides illustrate this:

| | |
|---|---|
| Ferric hydroxide . . . . . | $Fe_2(HO)_6$ |
| Ferrous hydroxide . . . . . | $Fe(HO)_2$ |

## CHLORINE AND THE CHLORIDES.

More than a hundred years ago — it was in 1774 — the Swedish chemist Scheele was studying the action of hydrochloric acid on a black powder known as the "black oxide of manganese." To his surprise a heavy greenish gas was produced. Sir Humphry Davy afterward called this gas *chlorine*.

Chlorine is a very suffocating substance, and in all experiments with it we must be careful to not breathe it. When we make the gas, and when we use it, there must be the utmost care to have our apparatus air-tight, and to prevent the escape of this noxious gas.

The following arrangement of our gas-making apparatus will enable chlorine to show several of its properties without escaping to poison the atmosphere of the room.

**Preparation of Chlorine.** — We will adopt the original method, and make the gas by the action of "black oxide of manganese" on hydrochloric acid.

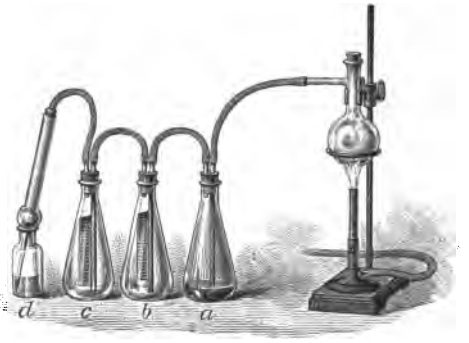


Fig. 58.

*Ex. 86.* — Into *a* I put 50 cc. water, and close the flask tightly with its stopper. By means of a piece of small wire I hang a strip of moist blue litmus-paper to the stopper of *b*, and press it down into the flask.

I fix a piece of "gold leaf" (Dutch metal) in the same way, and enclose it in *c*. I fill the bulb of a drying-tube loosely with cotton, and then fill the wide part of the tube with dry slaked lime and cork it tightly.

I then join all these parts, as shown in the cut, by means of rubber tubing, — the side-neck with the long glass tube of *a*, the short tube of *a* with the long tube of *b*, the short tube of *b* with the long tube of *c*, and the short tube of *c* with that of the drying-tube, the small end of which I place in a bottle, *d*.

Finally, to make the chlorine I put 25 g. of manganese dioxide, $MnO_2$, into the side-neck flask, pour in upon it 100 cc. strong hydrochloric acid, press home the solid rubber stopper, and apply a *small flame* of the Bunsen. The heat must be gentle. The chlorine will slowly drive the air over from flask to flask, and out at *d*; but if the joints are as tight as they may be, no chlorine will escape into the room.

The chlorine soon fills the side-neck flask, as may be seen by its color, and then bubbles through the water in *a*, and passes slowly on toward *d*.

What is the action of Cl and $H_2O$, shown in flask *a*?
What is the effect of Cl on the color of litmus, shown in *b*?
What is the effect of Cl on Dutch metal, shown in *c*?
Why are there no bubbles of Cl in the bottle *d*?

THE FACTS. — The chlorine comes from the hydrochloric acid. The reaction may be written as follows:

$$\underset{\text{Manganese dioxide}}{MnO_2} + \underset{\text{Hydrochloric acid}}{4HCl} = \underset{\text{Manganese chloride}}{MnCl_2} + \underset{\text{Water}}{2H_2O} + \underset{\text{Chlorine}}{2Cl}$$

Chlorine is a greenish-yellow gas, whose odor is pungent, and suffocating in the highest degree. Cold water dissolves about two and one-half times its own bulk of chlorine, and then has the same color as the gas itself. This

chlorine-water may be used instead of the gas for many purposes; it will cause the same chemical actions. We see also (flask *b*) that chlorine destroys the color of litmus, and (flask *c*) that it combines with "gold leaf," and that it is greedily absorbed by the lime in the "drying-tube."

The actions of chlorine on colors and on the metals are important. Let us study them further.

**Bleaching.** — Chlorine destroys the color of litmus (*b*). Will it also bleach other colors?

*Ex. 87.* — I fill a test-tube half-full of water, and color it with black writing-ink. I slip the rubber tubes off from the glass tubes of flask *a*, and then, carefully inverting the flask over the mouth of the test-tube, as in Fig. 57 I let a few drops of chlorine-water run into the black liquid.

*Ex. 88.* — I use cochineal, or an aniline dye, or the petal of a flower, or a piece of calico cloth, instead of the inky water in the last experiment.

*Ex. 89.* — I take a piece of paper covered with writing in red ink, and a piece of newspaper covered with printing, and moisten both with the chlorine-water.

Is the color discharged in both cases?

What is the coloring matter of printers' ink?

THE FACTS. — This power of chlorine to destroy coloring matters is made use of on a large scale in the art of *bleaching*. All colors which are made from vegetable or animal substances will be removed by chlorine *in the presence of water:* dry chlorine does not bleach. Nor are mineral colors (carbon of printers' ink) disturbed by it. The bleaching of linen and cotton goods and paper is done by the use of chlorine.

Chlorine also destroys bad odors, and it is much used in hospitals and elsewhere for this purpose. It is the most powerful *disinfectant*.

Chlorine destroys colors and odors, by taking hydrogen away from them. In some cases the color or the odor gives up its hydrogen directly to the chlorine, but in others it is *water* which gives up its hydrogen to the chlorine, and then the *oxygen* which is set free attacks the substance of the color or the odor and decomposes it.

Bleaching Powder. — Chlorine is absorbed quite greedily by slaked lime. The product is called *bleaching powder*, and this, instead of chlorine-water, or the gas itself, is used in the arts.

*Ex. 90.* — I remove a part of the lime from the drying-tube, Ex. 86, to a dish and cover it with water. After stirring it well, I pour it on a filter, and catch the filtrate in a test-tube. I then add a little of this clear solution to some inky water or litmus solution in another tube.

By passing chlorine-gas over dry lime in large chambers, instead of in a tube, immense quantities of this powder are made, to be used as a bleaching agent and disinfectant.

**The Chlorides.** — Chlorine readily combines with the metals and changes them into *chlorides*. This action took place in flask *c* (Ex. 86). The leaf of "Dutch metal" is said to be made of zinc and copper, and if so, the white residue is made up of zinc chloride and copper chloride. Let us search for the copper chloride.

But we should first examine some copper chloride itself, so that we may be able to identify it.

*Ex. 91.* — I put a little of a known specimen of copper chloride into a test-tube and add a little water. It easily dissolves.

What is the color of the strong solution?

What is the color if more and more water be added?

What happens if a piece of clean iron is left in the solution?

I next add to some solution in another tube a little ammonia, and then add more, and more, until the blue precipitate made at first is dissolved to a clear, fine blue liquid.

A compound of copper may be known by the green or blue *color* of its solution, by the *metallic copper* set free by iron, and by the action of *ammonia*.

*Ex. 92.* — I now carefully take the stopper and tubes from the flask *c*; the white residue of the metal is lifted out with them. I carefully wash this residue into a porcelain dish with as little water as will remove it. I look for the color, and if I find none I evaporate the solution down to a small bulk. Finally, I insert a bright clean needle or knife-blade and leave it a few minutes.

Compare these results with those in Ex. 91, and decide whether the chlorine produced copper chloride in the flask *c*, in Ex. 86.

CHLORIDES BY CHLORINE WATER. — Metals are changed to chlorides, not only by chlorine gas, but also by chlorine water. Thus:

*Ex. 93.* — I put a leaf of "Dutch metal," about two square inches, into a test-tube and pour upon it 5 cc. chlorine-water from *a*, and shake it until dissolved. I then carefully slide a clean iron nail down into the solution.

Does the color of the solution indicate a copper compound? Does the action of the iron prove its presence?

After half an hour take out the iron, and put in a fresh bright piece. If the action is over no copper will now be deposited. Keep this liquid for Ex. 97.

CHLORIDES BY HYDROCHLORIC ACID. — But the most usual way of changing metals to chlorides is by means of hydrochloric acid. Thus:

*Ex. 94.* — Into 5 cc. strong hydrochloric acid I put several

small bits of iron, such as small tacks. Hydrogen escapes with effervescence. I let the tube stand until the action is over.

Note the color of the iron chloride now in solution.

Get the solid chloride out, by evaporating the liquid.

The reaction may be written in this way:

$$\underset{\text{Iron}}{Fe} + \underset{\text{Acid}}{2\,HCl} = \underset{\text{Ferrous chloride}}{FeCl_2} + \underset{\text{Hydrogen}}{2\,H}$$

which shows that an atom of iron, Fe, takes the place of two atoms of hydrogen in the acid, and makes a new molecule of ferrous chloride, setting two atoms of hydrogen free. Thus the iron is changed into a chloride.

Iron has a stronger attraction for chlorine than hydrogen has, and can take the chlorine away from the hydrogen in the acid. This is also true of many other metals beside iron.

CHLORIDES BY AQUA REGIA. — When strong hydrochloric acid is mixed with nitric acid, they decompose each other; chlorine is set free, but stays in the solution. Now this mixture will do what neither of the acids separately can do: it will dissolve gold and platinum, changing them into chlorides. Gold has been called the king of metals, and this liquid which dissolves it is called *aqua regia*, the meaning of which is "royal water." Aqua regia is the most powerful of all agents for changing metals into their chlorides. For example:

*Ex. 95.* — I put 2 cc. strong hydrochloric acid into a test-tube and add about one-fourth as much strong nitric acid. Into this I drop two or three small "tacks." Chemical action is evident at once, and the iron slowly wastes away. The liquid changes color as the iron chloride increases, and toward the end orange vapors of a suffocating odor may appear. Keep this liquid.

**Two Chlorides of one Metal.** — The color of the liquid containing the iron chloride made by aqua regia is quite different from that of the iron chloride made by hydrochloric acid. Let us try to find the reason.

*Ex. 96.* — I repeat the experiment with hydrochloric acid (Ex. 94), and when the action is over compare the liquid with that made by aqua regia in Ex. 95.

Note the difference in color.

I next test a part of each of these two solutions separately with ammonia, just as I did the copper solution in Ex. 91.

Note the great difference in the two results.

Do the two solutions contain the same or different substances?

THE FACTS. — The iron chloride, made by hydrochloric acid, is green, while that made by aqua regia is yellow. When ammonia is added they behave very differently: the first gives a light green precipitate, while the last gives one that is reddish-brown. These differences lead us to think that there must be *two* different *chlorides of iron*, and this is true.

The green chloride is $FeCl_2$, called ferrous chloride.

The yellow chloride is $Fe_2Cl_6$, called ferric chloride.

In the names of the two chlorides of iron the endings *ic* and *ous* show the two proportions of chlorine, *ic* the larger and *ous* the smaller. With many other metals, also, chlorine forms both *ic* and *ous* chlorides.

*Ex. 97.* — When the iron, in Ex. 93, took the copper out of the copper chloride, what else occurred? To the liquid kept from that experiment add ammonia, and from the color of the result decide

Whether it contains copper chloride or iron chloride.

Whether it contains ferrous or ferric chloride.

You can then understand just what happened in Ex. 93. For if the ammonia gives the green precipitate, then ferrous chloride is present, and the iron must have changed the copper chloride into it. Thus:

$$\underset{\text{Copper chloride}}{CuCl_2} + \underset{\text{Iron}}{Fe} = \underset{\text{Ferrous chloride}}{FeCl_2} + \underset{\text{Copper}}{Cu}$$

Atoms of iron took the place of atoms of copper, and changed the copper chloride to the ferrous chloride.

**Hydrogen Chloride.** — This is another name for hydrochloric acid, — one of the strongest and most useful of all acids. It is made on a large scale from common salt by the help of sulphuric acid. On a small scale we may prepare it with the apparatus used for chlorine, leaving out only the drying-tube, as shown in Fig. 59.

*Ex. 98.* — I put 50 cc. strong sulphuric acid in *a*, press its stopper home, and join its long tube with the side-neck. I leave *b* empty (it should be dry), close it with the stopper and join to *a*. I put a strip of moist blue litmus-paper in *c*, and join it with *b*. The bottle *d* contains water, into which the tube from *c* is put with its end scarcely more than covered. Finally I put the materials, dry salt, 5 g., and strong sulphuric acid, 10 cc., into the side-neck, and close it tightly. The reaction sets in at once; but to keep it up, I apply a very gen-

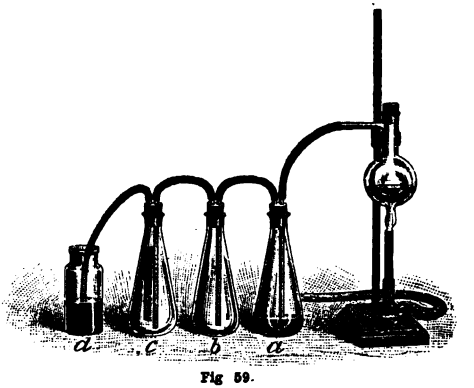


Fig 59.

tle heat so that bubbles will be seen in *a* a little faster than one can count. The acid in *a* is used to absorb moisture which comes over with the gas; the gas will be dried by the acid if it does not come over too fast.

Note the action of the hydrochloric acid on litmus.

Note the color of the gas.

What is the effect of the water in the bottle *d*?

*Ex. 99.* — I remove the cork and tubes from flask *b*, and then invert the flask with its mouth in a vessel of water. The water quickly rises to almost fill the flask.

How can you account for this result?

THE FACTS. — The reaction when sodium chloride, Na Cl, and sulphuric acid $H_2SO_4$, are gently heated is this:

$$NaCl + H_2SO_4 = HNaSO_4 + HCl$$

The hydrochloric acid, H Cl, is a colorless gas which instantly reddens litmus and dissolves with great freedom in water. The so-called hydrochloric acid, of commerce and the laboratory, is a solution of this gas in water.

**Composition of Hydrochloric Acid by Volume.** — If hydrogen and chlorine gases are mixed, they will combine when exposed to light, and, by measuring the gases, it has been found that it takes just as many cubic centimeters of hydrogen as of chlorine. Pure hydrochloric acid contains

1 volume of hydrogen and 1 volume of chlorine.

But now, how much hydrochloric acid will these two volumes of its elements make? It has been found that if 10 cc. of hydrogen and 10 cc. of chlorine are used, there will be just 20 cc. of the hydrogen chloride made.

One volume hydrogen and 1 volume chlorine make 2 volumes hydrogen chloride.

Now compare this fact with another noted under the

composition of water, p. 56, and still another noted under the composition of ammonia, p. 90.

We see that in the case of

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| H Cl, | 2 | volumes | of | the | elements | make | 2 | volumes of compound. |
| $H_2O$, | 3 | " | " | " | " | 2 | " | " |
| $H_3N$, | 4 | " | " | " | " | 2 | " | " |

Just *two volumes* of the compound gas is made every time! The same thing is found true of other compound gases also. Take the five compounds of nitrogen and oxygen, for example:

| | | | | | | |
|---|---|---|---|---|---|---|
| 2 volumes | N and | 1 volume | O | make | 2 | volumes nitrous oxide. |
| 1 volume | N " | 1 volume | O | " | 2 | " nitric oxide. |
| 2 volumes | N " | 3 volumes | O | " | 2 | " nitrous anhydride. |
| 1 volume | N " | 2 volumes | O | " | 2 | " nitrogen peroxide. |
| 2 volumes | N " | 5 volumes | O | " | 2 | " nitric anhydride. |

These are the results of experiments. And so, whether we have two volumes of the constituents, as in the second, or seven volumes, as in the fifth, there are just two volumes of the compound made when they combine. If we put all these facts into one statement, we have this *law:* The volume of a compound gas is TWO, whatever may be the number of volumes of the elements in it.

**Test for Chlorine and the Chlorides.** — Let a drop of silver nitrate solution be added to a solution of chlorine or of any chloride, and a white cloud or precipitate will appear. This white precipitate is silver chloride. It will become dark colored if left a little while in the sunlight. Try this test on several chlorides.

### THE CHLORINE GROUP.

There are three other elements which are so much like chlorine, that the four are together called the chlorine group. They behave so nearly like chlorine, and their

uses are so much less important, that we need not stop long with them in our study.

**Bromine.** — A little more than half a century after the discovery of chlorine — it was in the year 1826 — M. Balard, a French chemist, found another element, with properties much like those of chlorine. Its odor was found to be so strong that, in honor of this characteristic the element was called *bromine*. The word is from the Greek word βρῶμος (*bromos*), which means *stench*.

Bromine is not very abundant, but it does exist in the waters of some mineral springs, and in larger quantities in the waters of the sea. It is always found in combination, and its compounds are called *bromides*.

Bromine is a *liquid*, but a very gentle heat changes it to gas. It has a beautiful dark-red color.

This element readily combines with hydrogen and with metals; in this respect it is like chlorine, and like that element it is able to remove colors and to destroy bad odors.

**Iodine.** — M. Courtois was a chemical manufacturer in Paris. He was engaged in making soda, and was using the ashes of sea-weeds for this purpose. A dark-colored liquid was left in his kettles, and attacked the metal of which they were made. When some sulphuric acid was put with it, this liquid gave up a substance which, when heated, changed to a beautiful violet-colored vapor. It proved to be a new element, and it was called *iodine*. This, from the Greek word ἰώδης (*iodas*), means *violet-colored*.

Iodine is a constituent of sea-plants, as we may know from the story of its discovery. The sea contains it in small quantities, and so do the waters of some mineral springs. It is an element in sponges, and in oysters, and in some fishes. It is always found in nature, combined with other substances.

Iodine is a *solid.* A gentle heat melts it, and a little higher temperature changes it into its beautiful vapor.

Water dissolves it, but a single grain will take no less than 7000 grains of water to dissolve it. How small this proportion of iodine, and yet it gives to the whole body of the water a brownish-yellow color! Alcohol will dissolve it in large proportions. This solution in alcohol is the "tincture of iodine," used in medicine.

Iodine combines with hydrogen and with many metals. In this respect it is much like chlorine and bromine. Its compounds are called *iodides.*

Some of its compounds with the metals are remarkable for their very brilliant colors (Ex. 13).

But there is one effect of iodine which neither chlorine nor bromine can produce: it is the fine blue color it gives to starch. Thus:

*Ex. 100.* — I boil a bit of starch in a half test-tubeful of water, and after it has become cold I add a little iodine-water, made by vigorously shaking a crystal of iodine in water. Note the blue color produced.

*Compounds* of iodine do not blue the starch; the experiment may be made with potassium iodide. But if one of these can be decomposed the *free* iodine will show itself by the color; the experiment may be made by adding drops of nitric-acid to the solution of potassium iodide, and then adding the starch. The test for free iodine is starch.

The physician finds iodine very useful as a medicine.

The photographer finds it very valuable in the art of picture-making.

The element itself is sometimes used in medicine, but its compounds are more generally used in the arts.

**Fluorine.** — This element is found in combination with the metals, and all attempts to get it free have failed. Its chemical attractions are so strong, that if it be set free it

immediately combines with something else; even with the substance of the vessel in which the work is done. And yet it is remarkable that this is the only element which is not known to combine with oxygen. Its strongest affinities are for hydrogen and the metals. We cannot say whether fluorine is a gas, or would be if out of combination, as chlorine is. It is supposed that it would be, because it is a much lighter substance than chlorine. It is only nineteen times heavier, while chlorine is 35.5 times heavier than an equal bulk of hydrogen.

The mineral called *fluor spar*, $CaF_2$, is the most abundant compound of fluorine. When this is treated with sulphuric acid the fluorine leaves the calcium, Ca, and combines with hydrogen. This gives a gas,—the hydrofluoric acid, HF. This acid gas is useful in the art of etching glass; for wherever it touches glass the fluorine leaves the hydrogen to unite with the elements of the glass instead.

**Hydrogen Compounds.** — Fluorine, chlorine, bromine, and iodine behave toward hydrogen very much alike. Each one forms a single compound with that element, and each one combines with the hydrogen, atom for atom. Their names and formulas plainly show this fact. Thus:

| | |
|---|---|
| Hydrochloric acid . . . . . | HCl |
| Hydrobromic acid . . . . . | HBr |
| Hydroiodic acid . . . . . | HI |
| Hydrofluoric acid . . . . . | HF |

This is a very interesting fact, for there is no other non-metal, besides these four, that will combine with hydrogen *atom* for *atom*. In the case of oxygen, for example, the compound is water, $H_2O$, in which an atom of oxygen takes *two* atoms of hydrogen. An atom of oxygen refuses, absolutely, to ever take any less than two of hydrogen.

**General Behavior.** — These four elements behave very much alike in other respects. This is, to be sure, not quite so marked in the case of fluorine as in the other three, but in general they combine with the same elements and in the same proportions. And they are able to do the same kinds of work; as, for instance, chlorine bleaches colors; so does bromine, and iodine also, when in solution.

Chlorine is a *better* bleacher than bromine, — it is more powerful, and bromine is a better bleacher than iodine. But they are all bleachers.

**Atomic Weights and Properties.** — In general it is true in other cases, as well as in bleaching, that the chemical action of chlorine is more vigorous than that of bromine, and of bromine more vigorous than that of iodine. In fact, bromine will drive iodine out of combination and take its place, and chlorine will treat bromine in the same way. Now this order of chemical strength is just the order of their atomic weights, beginning with the smallest. Thus:

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| Atomic weight of | Cl | 35.5, | with | most | vigorous | chemical | action. |
| Atomic | " | " | Br | 80 | " | less | " | " | " |
| Atomic | " | " | I | 127 | " | least | " | " | " |

Their forms, gaseous, liquid, and solid, are in the same order; so are the densities of these elements. In fact, *the properties of these elements seem to depend, in some way, on their atomic weights.*

### *EXERCISES.*

**1.** *Study the actions of chlorides, bromides, and iodides on silver nitrate.*

*1.* Arrange three test-tubes, one with a little solution of sodium chloride, another with as much solution of potassium bromide, and another with as much solution of

potassium iodide. Add silver nitrate drop by drop, shaking the tube vigorously after each addition, until a drop fails to make a precipitate. Describe the precipitates. Note the effect of shaking them. Look carefully for some difference in their colors.

Then expose the tubes to sunlight, or for some time to diffuse light, and note the changes which occur in the colors.

*2.* Next test the solubility of these precipitates in ammonium hydrate. To do this, make a little fresh precipitate, and keep it from light as much as possible. When the precipitate has settled, *decant* the liquid, that is, pour it away carefully so as to leave the precipitate in the tube. Then pour upon it ammonium hydrate gradually, with shaking, until you can decide whether the precipitate dissolves. Do they all dissolve? Does one dissolve more easily than another?

*3.* Write a brief statement of all the facts which you have discovered, noting particularly the points of difference among the three compounds.

**2.** *Study the action of chlorides, bromides, and iodides on starch.*

*1.* Make a very thin starch-water, by boiling a minute piece of starch in considerable water, and cooling the liquid.

*2.* Dissolve a chloride, a bromide, and an iodide, each in water.

*3.* To a part of each solution add a few drops of nitric acid, and afterward add a little of the starch-water. There should be a difference in color produced; note it carefully.

To another part of each solution add a little of the cold starch-water, without the nitric acid. By this means you can decide whether nitric acid is necessary to bring out the colors observed before.

*4.* Add the starch-water to hot solutions instead of cold ones, and let them become cold. By this means you will learn the effect of heat.

*5.* Write a brief statement of all the facts which you have discoverd.

Chlorine, bromine, and iodine behave so very much alike, that it is not always easy to distinguish their compounds one from another. And then, too, there is another class of compounds called cyanides, which might be mistaken for some of these. Still the "silver-nitrate," and the "starch test," will help one to identify these three classes of compounds. In the case of the nitrate test, the pure precipitates differ a little in color, and more in solubility in ammonia, — the chloride being easily soluble, the bromide much less so, and the iodide almost not at all. In the starch test, nitric acid must be used to decompose the compounds, because it is only the *free* iodine which gives the blue, and the *free* bromine which gives the brown color.

**3.** *Take a few substances from the teacher, or a friend who knows what they are, and see if you can decide whether each is a chloride, bromide, or an iodide.*

## SULPHUR AND ITS COMPOUNDS.

**Native Sulphur.** — Sulphur seems to have been known and used in the most ancient days of which we have any account. The element itself, and its compounds also, are very abundant in the earth. The element itself, or *native* sulphur, is found in the neighborhood of volcanos, or where these fire-mountains have been active in time past. Large quantities are taken from mines in Sicily; in fact, a great part of the sulphur in commerce comes from this source.

**Native Sulphides.** — Sulphur is found *combined with metals* in the rocks and soils almost everywhere. These compounds are called *sulphides.*

The sulphide of iron is a good example. It is a brassy-looking substance, very common in many rocks. It is often found in the form of little cubes, as perfect as if they had been chiselled by an artist. Sometimes it is found in the form of thin, shining, yellow scales. It has been called "fool's gold," because the ignorant have been deceived by its color. It is iron disulphide, $FeS_2$, and is commonly called *iron pyrites.*

The sulphides are the substances from which the useful metals are often obtained. Lead, for example, is taken from the mineral called *galena*, but galena is the lead sulphide, Pb S. Silver, copper, and zinc are also among the useful metals, found in the rocks in combination with sulphur.

**Preparation of Sulphur.** — In its native state, as found in the mines of Sicily, sulphur is not combined with anything, but yet it is very far from being ready for the market. It is not in *combination*, but it is *mixed* with a great deal of earthy impurities from which it must be separated.

This is easily done, because sulphur is easily changed into vapor by heat, while the earthy impurities are not. Let the mixture be heated, then, and the sulphur-vapor will pass away and leave the impurities behind. The vapor, being cooled again, is sulphur, which is very much more free from earthy matter than before.

It is heated a second time to remove the impurities which the first process did not. This time the vapor is run into a large chamber, and it is condensed upon the cold walls in very fine powder. This powder is the "flowers of sulphur" which is so common.

When the chamber is smaller, its walls become too hot to collect the powder, which then melts, and the liquid runs down upon the floor. It runs into channels in the floor which lead it out of the chamber into moulds. In this way the familiar sticks of "roll brimstone" are made for the market.

**Effects of Heat on Sulphur.** *Ex. 101.* — I reduce a piece of brimstone to coarse powder and half fill a test-tube with it. I hold the tube in the hot air above the lamp-flame, and thus keep it from contact with too strong a heat, but lower it until I find the heat intense enough to melt the sulphur. The liquid should not be darkened, but be yellow and limpid.

*Ex. 102.* — I place the tube with the limpid yellow liquid where it will not be shaken, and watch the liquid while it slowly cools.

In what form does the sulphur solidify?

*Ex. 103.* — I now carefully re-melt the sulphur, and then make the yellow liquid a little hotter, — only a little.

Note the change in color.

I continue to heat the liquid gradually and incline the tube from time to time, and observe that the color deepens and the sulphur grows viscid until it refuses to flow at all.

I then heat it still more, and notice that the half-solid sulphur re-melts.

I continue to apply heat to find out whether the sulphur can be boiled, and what is the color of the vapor.

I then pour the liquid in a small stream into a vessel of cold water, and examine the sulphur after this sudden cooling.

THE FACTS. — Sulphur at ordinary temperature is a bright yellow solid, but when heated a little hotter than boiling water (114.5° C.) it melts to a limpid yellow liquid. Some care is required to preserve this yellow color, since it easily changes if the heat be much greater.

At a higher temperature, about 132°, the sulphur begins to grow viscid as well as dark colored. Darker and thicker it becomes as it gets hotter, until, at about 230°, it is almost black, and is so very viscid that it will not flow from the vessel even if turned bottom upward. Make the hot sulphur still hotter, and the dark-colored, almost solid substance grows less viscid again; it will flow in the vessel very much like thick syrup.

At a higher temperature, about 450°, the sulphur boils. The hot vapor is colorless, but if cooled a little it is condensed and becomes yellow.

PLASTIC SULPHUR. — If melted sulphur, just below its boiling-point, be poured into cold water, it cools into a substance which looks like India-rubber. If we handle it, we find that it is like India-rubber in other things besides its color: it is *tough* and *elastic*. This unusual form of sulphur is called *plastic sulphur*. It will not stay in this condition, but gradually it will become again yellow and brittle as at first. This plastic sulphur is almost as unlike the common form as if it were another element. In these two forms we find a good example of allotropism, p. 39.

CRYSTALS OF SULPHUR. — If melted sulphur, limpid and

yellow, be allowed to cool slowly (Ex. 102), fine needle-shaped crystals will be seen to shoot out from the walls of the vessel. They will increase in number and size, until finally, when cold, the solid mass of sulphur will be made up of these slender prisms interlaced and crowded together.

Crystals of sulphur are also found in nature, but their shape (rhombic octohedra) is quite unlike that of the artificial crystals obtained by fusion. In these two forms of sulphur we have a good example of *dimorphism*, — the property in virtue of which a substance may crystallize in two distinct forms. Sulphur is called a *dimorphorus* element, because it crystallizes in two shapes.

**Artificial Sulphides.** — A sulphide of copper can be made as follows:

*Ex. 104.* — I prepare a small coil of fine copper wire, No. 30, by winding the wire around a small lead-pencil or glass tube. I put a few fragments of sulphur into a tube, and then insert the coil and heat the sulphur to boiling. Watch for any sign of chemical action, and notice that, once well started, it will run to the upper end of the coil without the further aid of the flame.

Compare the substance of the coil now with the copper and the sulphur which were used.

*Ex. 105.* — The experiment may be made by mixing 4 g. flowers of sulphur with 8 g. of fine copper-filings, and heating this mixture in a test-tube.

THE FACTS. — When copper and sulphur are heated together they unite and form a compound which is grayish-black, and brittle. The lamp-flame brings them to the right temperature, when at once the action sets in. A vivid red glow begins, and goes quickly to the end of the coil or mixture. This glow is not due to the lamp-flame, because it goes on even if the flame is taken away. The

additional heat to make the glow must be caused by the chemical action itself. Thus:

| Materials. | | Products. |
|---|---|---|
| Cu + S | = | Cu S + heat |

Many other metals, like copper, will become sulphides when heated with sulphur. If a white-hot rod of iron is used to stir melted sulphur, the iron will become iron sulphide, and if zinc and sulphur are properly heated, zinc sulphide is obtained.

These artificial sulphides are not always the same as the native sulphides of the same metals. This artificial sulphide of iron, for example, is very different from the native pyrites. It is Fe S, while the native sulphide is $FeS_2$; the first is the ferrous sulphide, while the second is the ferric sulphide.

**Hydrogen Sulphide.** — The ferrous sulphide is very easily decomposed by acids.

*Ex. 106.* — I put a piece of ferrous sulphide,[1] not larger than a grain of wheat, into a test-tube, and pour upon it 1 cc. dilute hydrochloric acid (half water).

Notice the effervescence: what does it show?

Notice the odor in the tube.

Notice the color of the liquid when the action is over.

Add ammonia to the liquid and compare with Ex. 96.

THE FACTS. — By the action of the acid and sulphide a gas is set free with effervescence, and its bad odor — like that which arises from the putrefaction of certain things, such as eggs — proves this gas to be hydrogen sulphide. The ammonia test suggests the presence of ferrous chloride in the liquid. Hence the reaction is as follows:

$$FeS + 2HCl = FeCl_2 + H_2S$$

[1] This substance is sold in the form of "sticks," which is very much better than the powder or the granular form.

The $H_2S$ is the hydrogen sulphide: it is also called *sulphuretted hydrogen.*

This gas is found in the waters of "sulphur springs." All the sulphur which such springs contain is in the form of this gas. They owe their nauseous taste and odor, and their medicinal value, to hydrogen sulphide dissolved in their waters.

This gas is a powerful agent for changing the metals into sulphides; no other substance is so valuable to the chemist for this purpose.

**Preparation and Properties.** — In order to study the behavior of this gas without being annoyed by its escape into the room, we may employ our usual form of gas apparatus.

*Ex. 107.* — Put 40 cc. water into *a*, Fig. 60, to cover the lower end of the long tube. Add a cubic centimeter of solution of lead acetate, and a drop or two of hydrochloric acid to 40 cc. of water, and put the mixture into *b*. Into *c* put 40 cc. water containing some arsenious oxide solution, and several drops of acid, and into *d* put 20 cc. dilute ammonium hydrate (half water). Connect the parts of the apparatus, and *be sure that the joints are tight.*

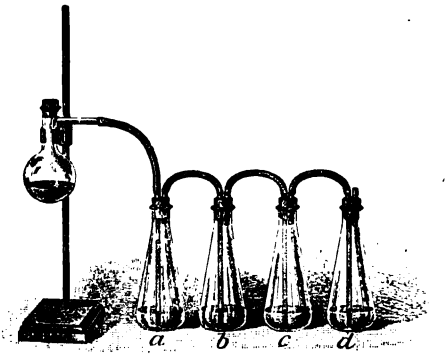


Fig. 60.

Finally, put 10 g. ferrous sulphide, in small pieces, into the side-neck flask, and pour upon it 30 cc. of dilute hydrochloric acid (half water), and stopper the flask at once.

The air will be gradually driven out of the flasks, and then the $H_2S$ will begin to show its presence by its action on the liquids through which it passes.

Note its color as seen above the liquids.
Note the results of its action on the lead compound.
Note its different action on the "arsenic."
Note the effect of the ammonium hydrate.

Let the action go on until the effervescence stops. Then take flask *a* out of the series, and transfer a little of its water to tubes for the following purposes.

What is the odor of this water?
What is the effect of adding to it drops of H Cl and then of lead acetate?
What is the effect of adding to it drops of H Cl and then of arsenic solution?
Does the water have the same effects as the gas itself?

THE FACTS. — Sulphuretted hydrogen is a colorless gas which is dissolved by water quite freely (*a*). The gas in solution behaves just like the gas alone, readily changing metallic compounds into sulphides.

This gas is also very soluble in ammonium hydroxide (*e*), but in this case the solution is not a simple one, as in water: there is chemical action between the two substances, and the result is *ammonium sulphide.*

**Use in Analysis.** — It is easy to see (flasks *b* and *c*) that, if one has a solution which contains either lead or arsenic, without knowing which, sulphuretted hydrogen will help him to decide, for lead sulphide is black while arsenious sulphide is yellow. So this gas is often used to *identify* the metals in the work of analysis.

But it is still more often used by the chemist to *separate* two or more metals whose compounds are mixed. An example will show how this may be done.

*Ex. 108.* — I mix 1 cc. strong solution of copper chloride and 1 cc. strong solution of zinc chloride, and then add 10 cc. $H_2S$ solution freshly made (Ex. 107). I let the black precipitate settle and then add drops more of the $H_2S$ solution, to see whether more precipitate forms. I must add the hydrogen sulphide *as long as it will make the precipitate*, but when it refuses to do more I filter the whole into a clean tube or bottle. I now have the black precipitate in the funnel, and the clear liquid in the vessel below.

This liquid should contain all the zinc and none of the copper. Test it by adding a few drops of ammonium hydroxide: the presence of zinc and absence of copper is shown by a *white* precipitate, without the *blue* which copper would yield.

The black precipitate should contain all the copper and none of the zinc. Test it by putting a little in a tube and warming it with a *few drops* of nitric acid to dissolve it, then adding a little water, and finally adding ammonium hydroxide. The copper will be shown by the familiar blue color.

### THE SULPHUR GROUP.

There are three other elements whose chemical actions are very much like those of sulphur. These are selenium, tellurium, and oxygen.

**Selenium.** — Selenium is a rare element, which is found in combination with some metals, such as copper and iron. These compounds are called *selenides*. The element is a solid. It burns in the air with a reddish-blue flame, and an odor peculiarly offensive, even worse than that of burning sulphur. It is a conductor of electricity, if it be first melted and then *slowly* cooled, but not a conductor if cooled quickly. Its symbol is Se.

**Tellurium.** — Tellurium is even more rarely found than selenium. The element is bluish-white, and has a fine luster. Like sulphur and selenium, it is found in the earth combined with metals, such as gold, silver, and lead. Its symbol is Te.

**Their Hydrogen Compounds.** — Sulphur, selenium, and tellurium unite with hydrogen, and in the same proportions. Each one combines with the hydrogen, not atom for atom, but one atom for two. In this respect they are like oxygen. Their names and formulas plainly show this fact. Thus:

| | |
|---|---|
| Hydrogen oxide . . . . . | $H_2O$ |
| Hydrogen sulphide . . . . . | $H_2S$ |
| Hydrogen selenide . . . . . | $H_2Se$ |
| Hydrogen telluride . . . . . | $H_2Te$ |

Now these four are the only non-metals which combine with hydrogen in this way — one atom with two. But these do so always.

**General Behavior.** — These four elements act very much alike in other respects. They combine with the same things and in the same proportions. This is not quite so well marked in the case of oxygen, which stands a little apart from the rest; but still the likeness is very strong, and these four form a well-marked natural group.

**Atomic Weights and Properties.** — The order of the atomic weights in this group is as follows:

| O | S | Se | Te |
|---|---|---|---|
| 16 | 32 | 79 | 125 |

This is also the order of their specific gravities. When equal volumes are weighed, O is found to be the lightest, and S, Se, and Te are heavier in this order. It is also the order of their melting-points; oxygen melts at a very low temperature, S melts at about 114° C., Se at 217°, and Te

at 500°. It is also the order of their energy in chemical action, that of O being greatest and of Te least. The gradually varying properties of this set of elements stand in the order of their atomic weights. This is the second case of this kind (see p. 151). Are there other instances? We shall see.

### SULPHUROUS OXIDE AND ACID.

**Burning of Sulphur.** — We have found (Ex. 8) that sulphur will burn freely in oxygen, with a rich blue flame, and that the product will dissolve in water, which will then redden blue litmus. We have also found (Ex. 43), that sulphur will readily burn in air. The product of combustion, in both cases, is sulphur dioxide, $SO_2$, which is more commonly called sulphurous oxide.

**Preparation of $SO_2$.** — This oxide is easily made by heating sulphuric acid with copper.

*Ex. 109.* — I put 10 g. of copper, in small pieces, such as clippings of sheet-copper or wire, into the side-neck flask

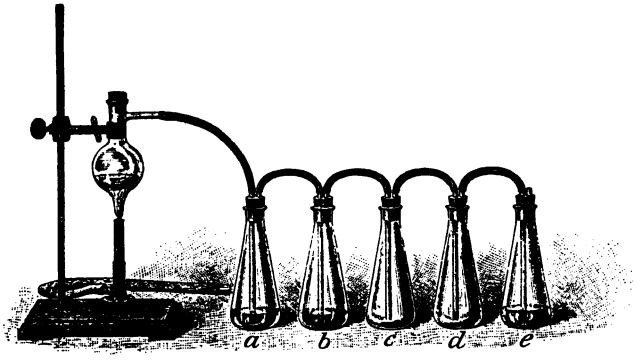



Fig. 61.

of the gas apparatus (Fig. 61), and pour upon it about 20 cc. of strong sulphuric acid.

I put 15 cc. of water in *a*, 15 cc. of sulphuric acid in *b*, and 50 cc. of water in *e*, and I then join the series *a*, *b*, *c*, *d*, *e*, as usual, the long tube of each with the short tube of the one in front of it, and *a* with the side-neck.

I now heat the side-neck flask with a small flame until effervescence begins, and afterwards just enough to keep the action going. The liquid will froth over if the heat is too strong. A little white vapor of sulphuric acid goes over into *a*, but the true appearance of the oxide should be seen in *b*, and in the other flasks.

What is the color of sulphurous oxide?

Can you decide whether the gas is absorbed by $H_2O$ in *e*?

**To discover the Properties of Sulphurous Oxide.** — When the action is over I at once take all the flasks apart, and then examine the gas they contain.

*Ex. 110.* — I thrust a strip of *moist* blue litmus-paper into *a* and leave it hanging, held by the stopper which I press in beside it. I suspend a strip of *dry* blue litmus-paper in the same way in *b*. Note any difference in the action of the moist and dry gas.

*Ex. 111.* — I lower the flame of a small taper, or of a splinter of wood, into the gas in *a*, and decide —

What is the effect of sulphurous oxide on fire?

*Ex. 112.* — If I boil a few small chips of logwood in water I get a rich wine-colored solution. I do this, and then, taking a test-tube half full of the colored water, I add some of the water from flask *e*.

What is the effect of $SO_2$ on this color?

*Ex. 113.* — I transfer a little of the water from *e* to a dish or tube, and test it to learn —

Whether it has the odor of the gas.

Whether it is an acid.

*Ex. 114.* — At this point I take the tubes from *c* and *d*, and, closing the flasks with solid corks, I stand them aside for use at another time (Exs. 117, 118). I pour the brown liquid carefully out of the side-neck flask, leaving the almost black sediment behind, and then put in about 50 cc. of water, shake it well, and filter the whole into a dish or bottle. I keep this deep-blue liquid also for use at a future time (Ex. 121). Pour the liquids from *a*, *b*, *e* into the waste, take out their stoppers and stand the flasks mouth down in water, and leave them until the noxious gas is absorbed.

THE FACTS. — Sulphurous oxide is a gas with no color, but possessing an odor of the most pungent and suffocating kind, — the odor of a burning "Lucifer match." The dry gas will not redden blue litmus, but when moist it will do so readily. It dissolves in water very freely, and this solution is an acid. It will also *bleach* the color of logwood and many other substances; this the dry gas will not do.

**Sulphurous Acid.** — The dry gas, and its solution, act quite differently on litmus, and also on coloring matter, and by this we know that the gas is not simply dissolved in the water, as in the case of chlorine, and of hydrogen sulphide, but that a chemical change has taken place.

The fact is, that the gas *unites* with water;

$$\underset{\text{Sulphurous oxide}}{SO_2} + \underset{\text{Water}}{H_2O} = \underset{\text{Sulphurous acid}}{H_2SO_3}$$

and while the oxide itself is not an acid, its compound with the elements of water is.

There are many other oxides like the sulphurous oxide in this respect, — they will combine with the elements of water, and thus form acids. All such oxides are called *anhydrides*.

This one is called the *sulphurous* anhydride, because it unites with the elements of water, and becomes *sulphurous* acid.

**Bleaching.** — The power of moist sulphurous oxide to remove colors makes this substance very valuable for bleaching. Articles of straw, silk, and wool are bleached on a large scale, by first moistening them and then hanging them in chambers in which sulphur has been burned. The sulphurous oxide does not act on the color itself, but it decomposes water and sets free a part of the hydrogen. This hydrogen then acts on the colors, and changes them to colorless compounds. Unfortunately these colorless compounds will be decomposed on exposure to air and light, and then the color returns.

Sulphurous oxide will not burn, nor will it allow the combustion of anything else; the fumes of burning sulphur will extinguish fire.

## SULPHURIC ACID AND THE SULPHATES.

**Some Properties of the Acid.** — If we examine a specimen of strong sulphuric acid, which can be found in every laboratory, we find that it is a heavy oily-acting liquid. For this reason it is called *oil of vitriol.* If we taste it — which we must not do without first diluting it with a large quantity of water — we find it to be sourer than the strongest vinegar. If we touch it, we find our fingers smarting almost as if it had been fire. If we drop it upon our garments, we find them turning red wherever touched by it, and that, some days after, the red spots crumble into holes.

When mixed with water the strong acid unites with it at once, and the mixture is heated as if by fire (Ex. 14). This violent attraction between the acid and water throws light upon some well-known facts, thus:

*Ex. 115.* — I put 2 or 3 cc. of oil of vitriol into a test-tube and place in it the end of a clean pine stick. After a few minutes I find the wood to be as black as if it had been scorched. Indeed, the wood is changed very much as it would be by fire. For by fire a black coal is left, while the other elements are driven away.

What are the elements in wood? See p. 101.
Which one of these is left behind by the acid?
Then which ones are taken out by it?
Why does the acid extract these?

One of our earliest experiments showed the singular effects of this acid on sugar (Ex. 2), and now you can no doubt find in the composition of the sugar, and this property of the acid, a good explanation of that action. Do so.

**Some uses of Sulphuric Acid.** — The chemist very often needs the gases, with which he works, to be perfectly dry; he remembers this strong attraction between sulphuric acid and the elements of water, and makes his gas bubble through some of the acid. It comes off dry.

Sulphuric acid is used in making a great many materials used in the arts, such as soap, soda, alum, and other kinds of chemicals.

It is used also in coloring cloth, in printing calico, and then, at other times, it assists in the work of bleaching.

Many *thousands of tons a week*, of sulphuric acid, are made and distributed over the world to be used for these and other purposes in the arts. The manufacture of sulphuric acid is one of the most important industries.

**Test for Sulphuric Acid.** — If, to any liquid which contains this acid, a solution of barium chloride is added, a white precipitate will be made, which cannot be dissolved by water or acids. Thus:

*Ex. 116.* — To half a test-tube full of water I add a drop of $H_2SO_4$, and a half cubic centimeter of HCl, and then I add a solution of barium chloride ($BaCl_2$) drop by drop. The white precipitate ($BaSO_4$) which comes, in spite of the hydrochloric acid, shows the presence of $H_2SO_4$. Try this test on a solution of any of the sulphates.

**The Manufacture of Sulphuric Acid.** — The process is founded on a few simple facts, which we can easily demonstrate, and for this purpose we have the flasks full of $SO_2$ kept over from Ex. 114. The $SO_2$ may be made by burning sulphur in oxygen (Ex. 8), or in air (Ex. 43), but we may use that already made in the other way. The question is now, How can this sulphurous oxide be changed into sulphuric acid?

*Ex. 117.* — I wet a shaving or splinter of wood with strong nitric acid and thrust it into the sulphurous oxide of flask *d* (Ex. 114), and leave it there awhile.

Notice the colored fumes which are produced.

What is the meaning of these? See Ex. 62.

It may be well to insert the splinter, wet with nitric acid, a second time. But finally I take it out, pour 10 cc. of water into the flask, and shake it well, to dissolve as much as possible of the contents, then pour it into a test-tube and label it *d*.

*Ex. 118.* — And now to find out whether the sulphurous oxide has been changed by the nitric acid I will dissolve the gas, which still remains in flask *c* (Ex. 114) in 10 cc. water, and then compare this solution with the other, labelled *d*, in this way: I first add a half cubic centimeter of hydrochloric acid, and then add drops of barium chloride, to each of my two solutions.

Note the difference in results.

What substance did the solution *d* contain?

**The Changes.** — When we bring sulphurous oxide and nitric acid together yellowish-red fumes appear (Ex. 117), and by these colored fumes we know that the nitric acid is being decomposed. Sulphurous oxide in water will yield no white precipitate with a mixture of hydrochloric acid and barium chloride, but after this action of nitric acid the solution will (Ex. 118). By this we know that the $SO_2$ has been so changed that its solution in water is sulphuric acid.

If now we could only take the water out of the sulphuric acid again, we could see what the $SO_2$ had been changed into. Unfortunately we cannot do this by experiment, but, fortunately, we can *represent* the process by formulas, as, in arithmetic or algebra, we often show by signs, the work which we do not actually do. From sulphuric acid, $H_2SO_4$, let us subtract water, $H_2O$.

$$H_2SO_4 - H_2O = SO_3$$

It must have been $SO_3$ which, when dissolved in water, gave the sulphuric acid which we found in *d*, and hence the $SO_2$ must have been changed into $SO_3$ by the nitric acid.

This $SO_3$ is called sulphur*ic* oxide. It is also called sulphuric anhydride, because it combines with water to make sulphuric acid.

The Facts. — What, then, are the facts which we have found? They are as follows:

1. The burning of sulphur yields sulphurous oxide.
2. Sulphurous oxide will take oxygen from nitric acid and become sulphuric oxide.
3. Sulphuric oxide dissolves in water and becomes sulphuric acid.

Application of These Facts. — These are the facts on which the manufacture of sulphuric acid is *founded.*

In the actual process there are a few others, but for the full details you may refer to some larger chemistry.[1]

The acid is made in immense lead-lined chambers. Such a chamber may be 100 feet long, 20 feet wide, and almost as high as it is wide. Sulphurous oxide goes over into this chamber from a furnace where sulphur, or iron pyrites, is burning. Nitric acid, made on the spot, also enters the chamber, while jets of steam are blown in, and a full supply of air is kept up.

The acid, which is made in this chamber, is too dilute for use, and it is made stronger by evaporation in lead pans, and then still stronger by evaporation in vessels of glass or platinum. These must be used instead of lead pans toward the last, because the strong acid will corrode the lead.

**The Sulphates.** — The salts made by sulphuric acid are called *sulphates*. We may study these as follows:

*Ex. 119.* — To make Zinc Sulphate. — I dilute some strong sulphuric acid by pouring 5 cc. of it into 40 cc. of water contained in a wide-mouth bottle, and drop into this small pieces of zinc. I let the action go on until the acid is used up, adding more zinc if necessary. When the effervescence has almost stopped, with zinc still left in the bottle, I filter the liquid to get rid of the black flakes which come from impure zinc. I next evaporate the liquid to one-half its bulk, and then let it cool. While it cools, crystals will be seen forming in the liquid. These crystals are zinc sulphate.

*Ex. 120.* — To make Ferrous Sulphate. — I make some dilute sulphuric acid, as in the last experiment, and

[1] In Roscoe and Schorlemmer, Vol. I. pp. 319–338, is a full account. In Cooley's Text-Book of Chemistry, p. 108, the reactions are given briefly.

drop into it small iron nails, such as small "tacks." I cover the bottle with a plate of glass or a square of heavy paper, and, after the action has gone on for some time, I lift the cover, and at the same time bring a match-flame to the mouth of the bottle and discover

What gas is set free by the action?

When the action is over I filter the liquid, and then evaporate it down to half its bulk, and let it cool. If the evaporation has gone far enough, crystals will appear. These crystals are ferrous sulphate.

*Ex. 121.* — To MAKE COPPER SULPHATE. — For this purpose the strong acid is needed instead of the dilute, and heat must be used. This work was done in experiment 109, and if the deep blue liquid has been kept (Ex. 114), we need not now repeat the experiment. Sulphurous oxide was set free instead of hydrogen, and the blue liquid contained the copper sulphate. Perhaps by this time blue crystals have made their appearance, but if not, I evaporate the liquid to smaller bulk and let it cool. The blue crystals are copper sulphate.

*Ex. 122.* — To PROVE THAT THE BLUE CRYSTALS ARE COPPER SULPHATE. — I dissolve a crystal in water and then add ammonia, as in Ex. 91, and also into another portion of the solution I put a piece of iron.

What are the proofs that the blue crystals are a compound of copper?

I dissolve another crystal and add a little hydrochloric acid and barium chloride (Ex. 116).

What is the proof that the blue crystals are a compound of sulphuric acid?

But a compound of copper and sulphuric acid must be the copper sulphate: the crystals are copper sulphate.

**Different Ways to make Sulphates.** — Many sulphates may be made by the action of the sulphuric acid on the metals directly. In the case of zinc and iron the action goes on in the cold, and yields hydrogen beside the sulphate. But in the case of copper, heat must be used to produce the sulphate, and sulphurous oxide and water, instead of hydrogen, are set free.

Silver, mercury, and some other metals are like copper in this respect. When heated with strong sulphuric acid they yield sulphates, sulphurous oxide, and water.

The sulphates may be made also by letting the acid act on the oxides or the hydrates of the metals, instead of on the metals themselves.

It is also found that this acid will sometimes yield *two salts of the same metal.*

It will do this with sodium, for when the acid is *gently* heated with common salt, as in Ex. 98, we get one salt of sodium, but when the heat is much stronger we get another. How can this be? Well, we see that the molecule of the acid, $H_2SO_4$, has two atoms of hydrogen, and if the heat be gentle an atom of sodium, Na, takes the place of only one of them, and we have $NaHSO_4$, but if the heat be strong both atoms of hydrogen are driven out by two of sodium, and we get $Na_2SO_4$.

The first, $NaHSO_4$, is the *acid* sodium sulphate.
The second, $Na_2SO_4$, is the *normal* sodium sulphate.

These two kinds of salts may come from any other acid in whose molecules there are two atoms of H, which may be driven out by a metal. All such acids are called *dibasic* acids. Sulphurous acid, $H_2SO_3$, and carbonic acid, $H_2CO_3$, are dibasic. They are dibasic, not because they have two atoms of H in a molecule, but because they have two atoms of H *which metals can displace.*

Acetic acid, $H_4C_2O_2$, is *mono*basic. Its molecule contains four atoms of H, but with metals it will give up only *one.*

An *acid salt* is one which still contains a part of the hydrogen which a metal can displace. A *normal salt* is one which contains none of the hydrogen which a metal can displace. The normal salts are generally neutral to litmus-paper, and are often called *neutral salts.*

### OTHER SULPHUR OXACIDS.

The sulphurous and sulphuric acids are the most important oxygen acids of sulphur; but besides these two there exists no less than six others. Look at their names and formulas.

| | | | |
|---|---|---|---|
| Hyposulphurous acid, | $H_2SO_2$; | Dithionic acid, | $H_2S_2O_6$ |
| Pyrosulphuric acid, | $H_2S_2O_7$; | Trithionic acid, | $H_2S_3O_6$ |
| Thiosulphuric acid, | $H_2S_2O_3$; | Tetrathionic acid, | $H_2S_4O_6$ |

Only one of these needs to be noticed any further at present, and that is the thiosulphuric acid. The sodium salt of this acid is very useful in photography: it dissolves away from the glass or paper the silver salts which have not been acted upon by light, and which, if left, would cause the picture to blacken. It is known as "hyposulphite of soda," but its true name is *sodium thiosulphate.*

### *EXERCISES.*

**1.** *Study the action of dilute acids on sulphides, sulphites, and sulphates.*

*1.* Put a little of the powder of some specimen of each of these compounds into a tube, moisten it with water, and add a little dilute hydrochloric acid. Watch for effervescence, or any other evidence of chemical action. Notice the odor of any gas which may be set free. If no action begins soon, heat may be used.

*2.* Use dilute sulphuric acid, making the experiments in the same way.

*3.* Write a brief statement of your results, pointing out the *differences* in the behavior of these three kinds of compounds.

**2.** *Study the action of barium chloride on sulphites and sulphates.*

*1.* Add drops of barium chloride to a solution of a sulphite and to a solution of a sulphate. Then compare the precipitates which appear.

*2.* Learn, by experiment, whether these two precipitates are alike soluble in hydrochloric acid.

*3.* See whether both these precipitates will appear if you add the hydrochloric acid to the solutions before you add the barium chloride. Compare Ex. 118.

*4.* Write a brief statement of your results, pointing out the difference in the behavior of the sulphites and the sulphates.

**3.** *Take from the teacher, or a friend who knows what they are, a few substances, and see if you can decide whether each is a sulphide, a sulphite, or a sulphate.*

## PHOSPHORUS, AND THE NITROGEN GROUP.

In 1669 a man by the name of Brandt, in Hamburg, was making experiments, hoping to find the "philosopher's stone," by the touch of which he would be able to turn any substance into gold. What he actually did find was a waxy-looking solid, yellow by daylight, but shining with a pearly white light *in the dark*. It burned furiously at the least provocation by warmth, and, on the whole, was so strange in its actions that the superstitious chose to name it "The Son of Satan." It proved to be an element, and it has since been known as phosphorus.

**Properties.** — Phosphorus comes to market in the form of round "sticks," four or five inches long, and about a half-inch in diameter. In color it is like yellow wax, but it is much harder than that substance, although soft enough to be easily cut with a knife.

Its most remarkable property is its strong attraction for oxygen. If brought into the air it begins at once to unite with oxygen, and waste away by a "slow combustion." But if gently heated at the same time, its combustion becomes rapid and furious.

To rub it with the warm fingers will inflame it. The friction of a knife-blade in cutting it will sometimes produce heat enough to set it on fire.

Such a substance can be safely handled only when it is under water. It is kept under water, and it ought to be cut under water, to avoid accident.

This element is a powerful poison.

**Red Phosphorus.** — Let some phosphorus be put into a vessel full of carbon dioxide or nitrogen, which will not act upon it, and then let heat be applied. In this

case the phosphorus will not burn, no matter how hot it becomes. But at a temperature of 240° C. a curious change takes place. "The melted phosphorus becomes solid, opaque, and of a deep red color." It is phosphorus in an allotropic form.

This "red phosphorus" may be exposed to the air and handled with very little danger.

The common and the red are not the only allotropic forms in which this extraordinary element exists. There is a white and flaky form, lately discovered by Remsen, and a black variety has been described by Thénard.

**Matches.** — Phosphorus is used, in large quantities, in making friction-matches. These are of several kinds, among which is the sulphur ("Lucifer") match, the paraffine match, and the safety match.

The sulphur match is made by dipping the end of a pine stick in melted sulphur, and then into a paste made of phosphorus and a little nitre, $KNO_3$, mixed in gum-water. Now, by rubbing the end of the match, we produce heat enough to set the phosphorus on fire; the burning of the phosphorus produces more heat, by which the sulphur is set on fire, and then the sulphur burning with still more heat sets fire to the wood.

The paraffine match is made in the same way, but paraffine is used instead of sulphur.

In this way the troublesome odor ($SO_2$) of burning sulphur is avoided. Sometimes potassium chlorate is used in the mixture, and then the match burns with a slight explosion.

The safety-match has no phosphorus in its tip; this element is spread on the side of the box instead. Red phosphorus is used. The match-stick is tipped with a mixture of sulphur or antimony sulphide, and potassium chlorate, and sometimes red lead or some other coloring substance is added.

To "light" this match, it must be rubbed against the phosphorus surface of the box.

**Phosphorus Oxides.** — There are two compounds of this element with oxygen.

| | |
|---|---|
| Phosphorus oxide . . . . . | $P_2O_3$ |
| Phosphoric oxide . . . . . | $P_2O_5$ |

The first is made when phosphorus is simply exposed to the air. The second is made by burning phosphorus in oxygen or air.

Let a small piece of phosphorus be hung, by a fine wire, inside a bottle; white vapors of $P_2O_3$ will fall from it and slowly fill the bottle. If then a little water is shaken in the bottle, some of these vapors are dissolved, and by adding a few drops of blue litmus solution, we prove the presence of an acid. Thus:

Phosphorus oxide . . $P_2O_3 + 3H_2O = 2H_3PO_3$ . . Phosphorus acid.

Let a small piece of dry phosphorus be placed on a little cup on a plate. Let it be touched with a hot wire, and immediately covered with a dry glass jar. Clouds of milk-white vapor are quickly formed, and, if everything is *dry*, snow-white flakes will soon be seen, some clinging to the walls of the jar, others falling like snow upon the plate. This snow-white solid is phosphoric oxide, $P_2O_5$.

Let a little water be poured on the plate, and the white solid will instantly dissolve with a hissing sound, like the sound of a hot iron in water. A little blue-litmus is reddened by this solution, proving the presence of an acid. Thus:

Phosphoric oxide . . $P_2O_5 + 3H_2O = 2H_3PO_4$ . . Phosphoric acid.

There is a third acid in this series; it is the hypophosphorus acid, $H_3PO_2$. But of these three we shall

notice further only the phosphoric acid, which is important on account of its useful salts, — the phosphates.

**The Phosphates.** — These are found in rocks and soils, in plants and in animals. It is in the form of phosphate that phosphorus exists most largely in nature. We find it more abundant in the seeds of plants than in other parts, and in the brain, the blood, and the bones of animals; and, accordingly, we find that the phosphates are much used to fertilize soils on which grain is to grow, and that they are also needful constituents in the food of man.

**Manufacture of Phosphorus.** — Phosphorus is itself extracted from bones. Almost half the weight of the bones in animals is *calcium phosphate*, $Ca_3 (PO_4)_2$, and about one-fourth of the weight of this phosphate is pure phosphorus.

If the skeleton of a man weighs 12 lbs. it contains about *one and a half pounds* of this element.

To obtain phosphorus from bones they are first burned; in this process they become very white and very brittle. They are afterwards crushed to powder; this powder is called *bone-ash*, and it is chiefly calcium phosphate.

From the bone-ash the phosphorus is obtained[1] by

1. Treating it with sulphuric acid.
2. Evaporating the solution to dryness.
3. Heating the residue nearly to redness.
4. Heating to redness with charcoal.

### ARSENIC.

The element arsenic is a solid substance with a steel-gray color, and a luster like the metals. It is very brittle and easily powdered. Its powder is sometimes sold under the false name "cobalt," to be used as a fly-poison.

[1] For details and explanations consult Roscoe and Schorlemmer, Vol. I. p. 460.

This element is sometimes found in the earth not combined with anything, but such *native* arsenic, as it is called, is not common. It is mostly found in combination with metals and sulphur.

In Silesia, Germany, there is found a mineral which contains arsenic with iron and sulphur; it is called *mispickel*, and its formula is Fe As S. This mineral is the source from which most of the arsenic of commerce is obtained.

The chief use of arsenic is in making shot. A small quantity of the element is melted with the lead. Lead alone is too soft; the arsenic hardens it.

**Arsenous Oxide.** — This is the principal compound of arsenic in commerce. Its formula is $As_2O_3$, and it is also called arsenic trioxide, because its molecule contains three atoms of oxygen. In the drug-stores it is often called white arsenic, but more generally simply "arsenic."

This oxide is made directly from arsenical pyrites, — another name for mispickel.

To get the oxide from mispickel the mineral is *roasted*, that is to say it is heated in a current of air. The hot oxygen of the air takes the arsenic from the hot mineral, and the two elements combine; arsenous oxide is the result.

This substance is a white solid; it can be dissolved in hot water, in cold water not so well. Its solution is almost tasteless and colorless, and without odor. It is a most fearful poison.

But while this compound is so fatal to the life of an animal, it has a strange power to prevent decay. It is used to destroy rats, mice, insects, and sometimes for the terrible purpose of taking the lives of men. On the other hand, it is made to do good service in the preservation of the stuffed or dried objects of natural history to be found in museums.

Arsenous oxide is an anhydride, that is to say, it be-

comes an acid by combining with the elements of water. By so doing it forms arsenous acid. Thus:

$$As_2O_3 + 3H_2O = 2H_3AsO_3$$

From this arsenous acid, $H_3AsO_3$, a large number of salts, — the arsenites, may be made. Sodium arsenite is very useful in calico-printing; it helps to "fix" the color, — in other words, prevent its fading. The common Paris-green is copper arsenite and copper acetate together.

**Arsenic Oxide.** — There is also arsenic oxide; its formula is $As_2O_5$. It is, like the other, an anhydride, for with water it yields arsenic acid, $H_3AsO_4$.

**Arsenic and Hydrogen.** — There is one compound of these two elements, a gas called *arsine*. It is made by the action of nascent hydrogen on a soluble compound of arsenic. It is very poisonous; Gehlen, who discovered it, lost his life by accidentally breathing a bubble of this gas. It burns freely, and, if made in the laboratory, it should be burned as fast as it forms.

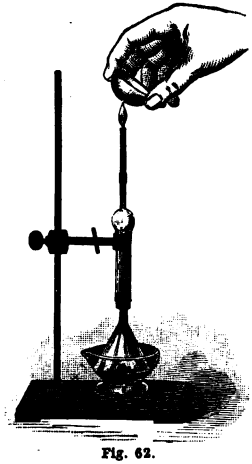


Fig. 62.

*Ex. 123.* — To make and burn this gas I use the simple hydrogen-flame apparatus of Ex. 20, shown again in Fig. 62. Before putting the zinc into the hydrochloric acid in the mortar to make the hydrogen, I dissolve a little arsenous oxide in water, by putting a very little of the powder — not more than will lie on the tip of a penknife-blade — into one or two cubic centimeters of water in a test-tube and boiling it.

I then drop several fragments of zinc into the dilute acid in the mortar, and at once lower the funnel to cover

them. Hydrogen is rapidly set free. It drives the air out of the apparatus, and when this is done I set fire to the jet. I now hold the bottom of a clean, dry porcelain dish right across the flame for a moment; no dark stain should be left upon it.

Then I pour the arsenous oxide solution into the mortar. Very soon the color of the flame becomes bluish-white. Again I press the cold porcelain for a moment down into the flame; a lustrous brown stain is left upon it. I let the flame play on other parts of the dish, touching it here and there on the inside also. A large surface may be thus covered with the brown arsenical mirror.

At the moment when hydrogen is set free by the zinc it attacks the arsenous oxide and changes the arsenic into arsine, $AsH_3$, and its oxygen into water.

The arsine burns with a livid flame, thus:

$$2\,AsH_3 + 6\,O = As_2O_3 + 3\,H_2O$$

arsenous oxide and water being the products of the action. But the cold porcelain cools the flame, and then only the hydrogen of the arsine burns, while the *arsenic is deposited* on the dish in the form of a lustrous mirror. This mirror will appear, even when the quantity of arsenic used in the experiment is astonishingly small. In fact, this experiment is a most excellent way of testing for arsenic in any suspicious substance. It is known as *Marsh's test.*

The chemist has studied the compounds of arsenic, until he has so well learned their characters, that he can tell with great certainty whether they are present in a substance or not, and Marsh's test is his most trusted method. It is true that antimony will also give a stain on porcelain in the same way, but its color is velvety black, and there are other well-marked differences which are well known to every practical chemist.

### THE NITROGEN GROUP.

PHOSPHORUS and arsenic are much alike in their chemical actions, and in some things both resemble nitrogen: these three are the members of the nitrogen group of non-metals.

**Their Hydrogen Compounds.** — These elements unite with hydrogen, and in the same proportions. The names and formulas of their compounds plainly show this fact. We have

| | |
|---|---|
| Ammonia (amine) . . . . . | $H_3N$ |
| Phosphine . . . . . | $H_3P$ |
| Arsine . . . . . | $H_3As$ |

We see that, in this group, each element unites with hydrogen, one atom for three. In other respects the resemblance among them is not so striking as it is in the chlorine group or in the sulphur group.

The order of their atomic weights is as follows:

| N | P | As |
|---|---|---|
| 14 | 31 | 75 |

This is also the order of their densities. The strength of their attraction for hydrogen is less, as the atomic weight is greater, in the same order. But among these elements the relation of properties to atomic weights is not so close as in the cases of the chlorine and the sulphur groups.

## SILICON, AND THE CARBON GROUP.

**Silicon.** — The element silicon is in many respects very much like carbon.

It is usually a dark-colored powder, but it has two other forms, which are a little like graphite and diamond.

Silicon, next to oxygen, is the most abundant element in the world, but it is never found alone; it is combined with oxygen, and the two are not easily separated. Immense quantities of silicon, in this condition, are hidden in the sandstone rocks, so very common, yet very few persons who are not chemists have ever seen the element itself.

**Its Oxide.** — The compound of silicon and oxygen is known as *silica*. Its true chemical name is silicon dioxide; its formula is $SiO_2$. Common sand is chiefly silica, and the sandstone rocks are masses of silica, mixed with many impurities to be sure, but chiefly silica. And besides, there are many purer forms, such as the following:

*Flint* is a kind of very hard stone which is sometimes white, sometimes brown or black, but it is always silica, and a much purer form than sandstone. In some places very fine transparent diamond-shaped crystals are found; they are called *rock-crystal*, and consist of pure silica. The common name of these hard varieties of silica is *quartz*.

The beautiful *amethyst* is quartz crystal, with a delicate purple color.

The precious *opal*, the *crysoprase*, and the *bloodstone* are little else than silica.

*Jasper* is a very fine-grained form of silica, colored generally red, but sometimes black.

The *agate* is a form of silica in which many tints of color are delicately arranged in stripes or bands. When the colors are few and very regularly arranged the agate is called an *onyx*. When the color is uniform, and a pearly white, the stone is a *chalcedony*, but when red it is called *carnelian*.

### THE CARBON GROUP.

Silicon behaves very much like carbon in its chemical actions. The two are much more like each other than like any other non-metals, and they together are called the *carbon group*.

**Their Hydrogen Compounds.** — These elements unite with hydrogen in the same proportions. They yield

Methane (carbon hydride) . . . . . $CH_4$
Silicon hydride . . . . . $SiH_4$

We see that, in this group, each element unites with hydrogen, one atom with four.

**Their Oxygen Compounds.** — They are not only alike in their combination with hydrogen, but also with oxygen, for we have carbon dioxide, $CO_2$, and silicon dioxide, $SiO_2$. And then, just as from carbon dioxide we get carbonic acid, by its union with water, thus:

$$\underset{\text{The dioxide}}{CO_2} + \underset{\text{Water}}{H_2O} = \underset{\text{The acid}}{H_2CO_3}$$

so from silicon dioxide we get silicic acid,

$$\underset{\text{The dioxide}}{SiO_2} + \underset{\text{Water}}{H_2O} = \underset{\text{The acid}}{H_2SiO_3}$$

The likeness does not stop here, for just as carbonic acid yields salts — the carbonates, so silicic acid yields salts — the silicates.

**The Silicates.** — These salts are very abundant in na-

ture, and very useful in the arts. The natural silicates occur in soils: clay is one of them. They also make up large rock masses: the slates are of this kind. Clay and the slate rocks alike are made chiefly of aluminum silicate.

Artificial silicates are easily made. It is only necessary to melt together some silica, — say fine white sand, and potassium, or sodium, or calcium carbonate, to produce the potassium, sodium, or calcium silicate. These artificial silicates are called *glass.*

In each different kind of glass there are at least two kinds of silicate. In common "window glass" there are sodium and calcium silicates; in "flint glass" there are potassium and lead silicates. Most ornamental glassware, such as vases, fine goblets, and decanters, are made of flint glass.[1]

## BORON.

**The Element.** — The element boron can be obtained in two forms; as a dark-brown powder and as fine crystals, almost as hard as diamond. It is never found free in nature, but some of its compounds are quite abundant. The most interesting are boric acid and borax.

**Borax.** — There is a lake in California (Borax Lake) whose waters hold a large quantity of borax in solution, and borax for the market is obtained by evaporating this water. It comes out in the form of a mass of white crystals. The chemical name of this substance is sodium biborate. But the pure and dry biborate is not quite the same as the crystals, for they hold a large portion of water in combination, while the biborate has none.

This is shown by the formulas, —

| | |
|---|---|
| Sodium biborate . . . . . | $Na_2B_4O_7$ |
| Borax crystals . . . . . | $Na_2B_4O_7 + 10H_2O$ |

[1] Roscoe and Schorlemmer, Vol. II., Pt. I., pp. 462 to 490.

Now, this water is a part of the crystal, for if we drive it away by heat the borax is no longer crystalline. This water is, therefore, called the *water of crystallization.* The crystals of a great many other things also hold water of crystallization.

*Ex. 124.* — I make a solution of 5 g. of borax in 20 cc. of hot water in a porcelain dish, and then add to it, in small portions, 1 cc. strong sulphuric acid. I let this stand until cold. Then notice and describe the crystals which form. Are they crystals of borax?

*Ex. 125.* — To answer this question I put some of the crystals into one porcelain dish, and as much borax into another, and pour upon each 10 cc. of strong alcohol. I stir them well, and then fire them both with a match-flame, and notice the color of the flames, especially around the edges.

What evidence that the crystals are not borax?

**Boric Acid.** — Borax is changed to boric acid by the action of strong sulphuric acid. Boric acid when dry appears in the form of glistening scale-like crystals (Ex. 124). These dissolve in alcohol, and tinge the flame of alcohol with green (Ex. 125).

Boric acid is sometimes made on a large scale in this way, that is by the action of a strong acid on borax. But it is also obtained by evaporating natural waters which hold it in solution. Such water is found in Tuscany. It is in a volcanic region, and jets of volcanic steam, out of the earth, are directed into this water. The water evaporated by the volcanic heat leaves boric acid.

**No Hydrogen Compound.** — Boron is the only non-metal which does not combine with hydrogen. We have seen that elements, which combine with hydrogen in the same way, are very much alike also in most of their chemical actions. Now here is an element which seems to have

no disposition at all to unite with hydrogen, and, curiously enough, we find it standing apart from all the other non-metals in other respects too. When free, it is a little like carbon and silicon. In some of its properties it resembles the nitrogen group; in others it is more like the metals than the non-metals.

Boron does not fairly belong to either group of non-metals, but on the whole it stands nearer to the nitrogen group than to others, and it is usually named in the list of that group.

But, in both its forms, it differs from the members of that group. For example: the crystallized variety will not readily combine with oxygen, while the elements of the nitrogen group, except nitrogen itself, will; and the other variety — the brown powder — when heated to redness, will combine with nitrogen; this the other members of the nitrogen group will not do.

## VALENCE.

We have seen that every one of the non-metals, except boron, will combine with hydrogen. But we have found that they do not all combine with it in the same proportions. In some cases they combine with it, atom for atom (p. 150), in others one atom for two (p. 162), and in other cases the proportions are still different (pp. 182, 184). What is the meaning of this?

**A Difference in Atoms.** — The proportions of the elements, in these compounds of the non-metals with hydrogen, are found by actual analysis, and then they may be shown by their formulas, each symbol of an element standing for one combining weight. Let us take one formula from each of the four groups which we have studied, as an example, —

$$HCl \qquad H_2O \qquad H_3N \qquad H_4C$$

and we see that one combining weight of chlorine combines with *one* of hydrogen, but that one of oxygen is not satisfied with so little; it must have *two*. One combining weight of nitrogen refuses to unite with less than *three* of hydrogen, while one of carbon demands *four*.

But we may as well say atoms, as combining weights, for, by the atomic theory, a combining weight of an element represents one atom. If so, the formulas show that one atom of chlorine can hold one atom of hydrogen, that an atom of oxygen is able to hold two, while an atom of nitrogen is able to hold three, and an atom of carbon, still more powerful, is able to hold four. In this way it appears that these atoms differ in their power to hold atoms of hydrogen in combination in a molecule.

VALENCE. — Now this power of an atom to hold a definite number of other atoms in a molecule is called its *valence.*[1]

The valence of an atom is *measured* by the number of hydrogen atoms which it can hold in combination. Thus the valence of chlorine is 1, of oxygen is 2, of nitrogen is 3, and of carbon is 4.

The valence of an atom is *described* by a prefix. Thus chlorine is said to be *uni*valent, oxygen *bi*valent, nitrogen *tri*valent, and carbon *quadri*valent.

The valence of an atom is *shown to the eye* usually by primes or dashes, written with the symbol, thus:

Cl′ or Cl– shows that chlorine is a univalent element.
O″ " O= " " oxygen " bivalent "
N‴ " N≡ " " nitrogen " trivalent "
C⁗ " C≣ " " carbon " quadrivalent "

The place of the dashes is of no consequence; they may be written above the symbol, below, at the right, or at the left of it. Thus:

Cl— —O— —N— —C—

It is the *number* of dashes which shows the valence of the element.

**Substitution is Governed by Valence.** — Chlorine will take the place of hydrogen in methane. To bring this about it is only necessary to mix the two gases and put the mixture in diffuse light. When this is done, it is found that the chlorine will take the place of the hydrogen *gradually,* until none is left. Methane, $CH_4$, may as well be written C H H H H. Then it is found that, —

[1] This property is also called quantivalence, and equivalence, and valency, by different writers. To call it valence is one of the later suggestions, and one worthy of general adoption.

| | | | | |
|---|---|---|---|---|
| C H H H H | is first | changed | to | C H H H Cl |
| C H H H Cl | " then | " | " | C H H Cl Cl |
| C H H Cl Cl | " " | " | " | C H Cl Cl Cl |
| C H Cl Cl Cl | " " | " | " | C Cl Cl Cl Cl |

Notice that the substitution of Cl for H goes on gradually, *atom for atom.* The carbon gives one atom of hydrogen for one of chlorine every time.[1]

The chlorine and hydrogen atoms have the same valence. Like bronze and copper pennies, they are different kinds of matter, but have the same value in making change. One atom of chlorine may be exchanged for one atom of hydrogen, and no more, in any chemical action. In fact, all univalent substances displace one another *atom for atom.*

But a bivalent atom is worth as much as two univalent atoms, and, in chemical action, the exchange must be made *one atom for two.* So one quadrivalent atom is worth two that are bivalent, and two trivalent atoms are worth three that are bivalent, in all chemical changes.

**What is the Valence of Boron?** — Boron will not combine with hydrogen, but it will with chlorine. The chloride is $BCl_3$. Here we see one atom of boron holding *three* univalent atoms of chlorine, which shows that boron is trivalent.

In this way the valence of many other elements has been found; if they have no hydrogen compounds their chlorine compounds, or their compounds with some other element whose valence is known, may be used instead. The valence of the metals has been found in this way.

**Valence Useful in Writing Reactions.** — The valence of the elements tells us how many atoms must take part in a reaction, and helps one to see how many molecules of

[1] The H which is driven out combines with other atoms of Cl to form H Cl.

each substance are involved in the change which goes on. For example:

We have found that sodium with hydrochloric acid yields sodium chloride and hydrogen (Ex. 76). But sodium is univalent, and so we write

$$Na' + H'Cl' = NaCl + H$$

So also zinc and hydrochloric acid yield zinc chloride and hydrogen (Ex. 78). Shall we write

$$Zn + HCl = ZnCl + H$$

No, because zinc is bivalent; one atom of zinc is worth *two* atoms of hydrogen, and *must* take the place of two. But there is only one in one molecule of H Cl, and therefore we must write

$$Zn'' + 2\,H'Cl' = Zn''\,Cl'_2 + 2\,H$$

Many reactions already studied will now seem clearer in the light of this explanation, furnished by valence, and in the future study of chemical actions valence will help us much.

**The Valence of an Element changes.** — But the valence of the same element is not always the same. It depends partly on the other substance with which it acts. Sulphur is bivalent with hydrogen, $H_2S''$, but it is quadrivalent with oxygen $S^{iv}O''_2$. And nitrogen, a trivalent element, sometimes has a valence of five. Still the changes in valence are very regular, and they do not greatly hinder our using valence as a guide in writing common reactions.

## METALS.

**What is a Metal?** — We think of a metal as a substance which is heavy, hard, and lustrous, and a good conductor of electricity and heat, because most of the common metals have these properties in high degree.

But there are some metals which are as soft as wax and lighter than water. In some conditions, metals are not at all lustrous, and some are not very good conductors. We must look further than to these properties for a real difference between metals and non-metals. It will be found in their chemical actions.

The compounds of other non-metals with hydrogen and oxygen are *acids*, while the compounds of metals with hydrogen and oxygen are *bases*. Remember that acids and bases are just opposite in character, and yet both contain hydrogen and oxygen. The *other element* which is combined with these two must be very different, in order to make an acid in one case and a base in another. And those which form the acids are the non-metals, while those which form the bases are the metals.

*A metal is an element whose compound with hydrogen and oxygen is a base.*

There is another chemical action in which metals and non-metals differ. Remember what has been said about salts, that they are made by putting another element into an acid in place of the hydrogen. Now, these elements which can take the place of the hydrogen in acids are the same ones which, united with hydrogen and oxygen, form the bases; they are the metals. This gives us another definition:

*A metal is an element which will take the place of hydrogen in an acid and form a salt.*

Accordingly the chemist divides the seventy-one elements into two classes, metals and non-metals, and in a general way the division is right. But nature does not draw any such sharp line of division through the list of elements. There are some elements whose compounds with hydrogen and oxygen are sometimes bases and sometimes acids. This is true of iron. The fact is, that while some of the elements are perfect non-metals, and some are perfect metals, there lie, between these, others which are less and less perfect, and some which are almost as much one as the other. In nature there is a gradual difference, in the properties of the elements, running from one end of the list to the other.

**Number and Abundance of the Metals.** — We have seen that the non-metals are included in four groups, and that they number only fifteen. All others of the seventy-one elements are metals.

Most of this large number of metals are rare substances seldom seen or little used. Not more than fifteen or twenty are abundant in nature or useful in the arts.

In our study of the metals we will select those which will teach us the most chemistry, are the most abundant, and the most useful. The study of the complete list of metals and their compounds need be attempted only by students who are to become chemists. To such this present course is simply a preparation.

**Occurrence of Metals in Nature.** — There are a few metals which are sometimes found free in the earth, not pure, but simply mixed with other things. Gold and silver and copper are examples. Metals when found in this condition, uncombined with other elements, are called *native metals.*

But the metals are usually found in combination with non-metals. In fact, the solid earth is made up of such compounds. But the rocks generally contain so much besides these compounds, or it is so difficult to get the metals out of them, that they are worthless for this purpose.

There are, however, some parts of the rocky masses, which are made up of metallic compounds, so rich in metal that they are valuable substances from which to get the metals themselves. Such compounds of the metals are called *ores*.

The work of taking these ores out of the earth is *mining*, and that of getting the metal out of the ore is *metallurgy*. Mining is a mechanical operation, and we will not stop to describe it, but metallurgy is a chemical art, and we will study it in connection with several of the common and the useful metals when we reach them, as a good example of the application of chemistry to useful purposes.

The most valuable ores of the metals are the oxides, the sulphides, the chlorides, and the carbonates.

## THE POTASSIUM GROUP.

### POTASSIUM. *K*.

**Description of the Metal.** — Most metals are very hard, but potassium is as soft as wax. It is easily moulded by the fingers or cut with a knife. When freshly cut, the surface shines with a blue-white luster.

Most metals are heavy, but this one is lighter than water. A piece dropped on water will float like cork. Nor is this the only thing that will surprise one, who for the first time drops potassium on water; the metal will instantly take fire. Violet-colored flames will burst from it, while the melted globule will run wildly over the water, wasting away all the time, until, when nearly all gone, it will usually put a stop to the action by a small explosion.

**Its Chemical Action on Water.** — The product of this action is found to be a base. (How would you prove it?) It is the potassium hydroxide. We may write the reaction

| HHO | + | K | = | KHO | + | H |
|---|---|---|---|---|---|---|
| Water | | Potassium[1] | | Potassium hydroxide | | Hydrogen |

The atom of potassium is univalent. It takes the place of one of the two atoms of hydrogen in a molecule of water, and forms a molecule of the hydroxide.

Hydrogen and much heat are also set free; in fact, heat enough to set the hydrogen on fire. A little of the potassium itself also burns, and it is this which gives violet color to the flame.

**Occurrence in Nature.** — Compounds of this metal are present in all fertile soils, and from the soil they pass

[1] The Latin name of potassium is *Kalium*, and the symbol, K, is taken from this.

into the bodies of plants. In some places potassium nitrate is found in the earth in large quantities; this is true of the dry tropical countries like Egypt and India. It is very soluble, and can be easily washed out by water. It is known in commerce as *nitre* or *saltpetre*. Potassium compounds are obtained for commerce either from plants or from the soils which contain them in large enough quantities.

**Potassium Carbonate.** $K_2CO_3$.—When wood is burned the compounds of potassium in it are all changed to potassium carbonate, which becomes a part of the ash. To get it out, the ashes are washed with hot water, for it is well known that the carbonate is soluble, and then, to get it in solid form, the solution is evaporated until the carbonate crystallizes out. Of course everything else in the ashes, which is soluble, will come out with the carbonate. The product is very impure; it is called *potash*.

The only *chemical* action in this process is the burning of the wood. The method of getting the potash out of the ashes may always be used to separate a substance which is soluble when it is mixed with others which are not.

Many of the other useful compounds of potassium are made from the carbonate by chemical processes.

**Potassium Hydroxide.** KHO.—Potassium will drive the metal calcium out of combination with hydrogen and oxygen, and take its place. This fact being known, the chemist uses it to get potassium hydroxide.

He mixes slaked lime, $Ca''H_2O_2$, with potassium carbonate in water, and boils the mixture. The metals simply change places.

| $Ca''H_2O_2$ | + | $K'_2CO_3$ | = | $CaCO_3$ | + | $2KHO$ |
|---|---|---|---|---|---|---|
| Calcium Hydroxide | with | Potassium Carbonate | become | Calcium Carbonate | and | Potassium Hydroxide |

The calcium carbonate is a white solid which falls to the bottom, while the potassium hydroxide stays in solution. This liquid is then boiled down in iron pans.

Potassium hydroxide is the same as caustic potash. It has a remarkable attraction for water. It cannot be kept except in tight bottles, for if left in the open air it will take moisture enough to completely dissolve it. Some other things have this same property; they dissolve in water which they get out of the air. Such substances are said to be *deliquescent.*

**Other Compounds made from $K_2CO_3$.** — There are a great many compounds of potassium which can be made by starting with the $K_2CO_3$.

*Ex. 126.* — Make a little potassium chloride. To do this, first dissolve $K_2CO_3$, as much as you can, in say 50 cc. of water. Then add HCl slowly, until, after shaking it, the liquid will redden a bit of blue litmus-paper. Finally evaporate the liquid to a small bulk and let it cool. Pour the liquid away from the crystals, and dry them on filter-paper. Keep this potassium chloride, KCl.

Why was the litmus-paper used?

What reaction took place?

*Ex. 127.* — Make some potassium nitrate in the same way, only use nitric instead of hydrochloric acid. If the evaporation is carried far enough, you can watch the crystals growing in the liquid while it cools. Keep this $KNO_3$.

*Ex. 128.* — Make a little acid potassium tartrate. To do this first make a *strong* solution of $K_2CO_3$ in one tube or bottle, and a strong solution of tartaric acid in another. Then add the acid to the carbonate, until a piece of litmus-paper is reddened by the mixture. A white crystalline precipitate, or solid, will be made. This is the acid potassium tartrate, known in the shops as "cream of tartar." Keep this.

When the two *liquids* were put together in the last experiment a *solid* made its appearance. Any solid made in this way is called a *precipitate.* It takes the solid form because it is not soluble in the liquids. This, and one or two others, are the only precipitates which we can make from potassium compounds, because the acid tartrate and one or two others are the only potassium compounds which do not readily dissolve. Even these dissolve a little, and on this account no precipitate will come unless the solutions used are strong.

**Flame Color.** — The flame of potassium is violet, and this color is seen when any compound of this metal is decomposed by heat. Thus:

*Ex. 129.* — I take a piece of platinum-wire and bend one end into a round loop about as large as this O. I moisten this loop and plunge it into the K Cl made in Ex. 126. A little of the salt will cling to the wire, and I hold it in the mantle of a colorless flame (Fig. 63), and notice the violet color of the flame above.

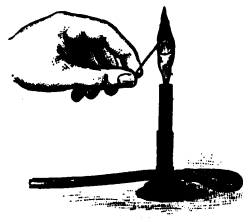


Fig. 63.

Look at this colored flame through a piece of cobalt-blue glass. Is the violet visible?

*Ex. 130.* — Thoroughly clean the loop so that it will not color the flame, and then try the $KNO_3$ of Ex. 127 in the same way.

Try also the acid tartrate of Ex. 128.

The color will sometimes come more surely if the loop is moistened with hydrochloric acid. Try $K_2CO_3$ without hydrochloric acid and then with it.

None but potassium compounds will give this color.

### SODIUM. *Na.*

The metal sodium is so much like potassium that a separate description is scarcely needed. Moreover, the student has already seen or handled it in several experiments of this course, and can remember its appearance and actions.

**Sodium in Nature.** — Common salt is made of sodium and chlorine, it is sodium chloride, Na Cl. Immense quantities of this are in the seas, and large beds of it are found in the earth. It is also in the salt-springs which are found here and there. Many other compounds of sodium also occur, in large quantities, in the earth; the carbonate and the nitrate may be mentioned.

**Sodium Carbonate.** $Na_2CO_3$. — This, next to common salt, is the most important compound of the metal. It may be made from the ashes of sea-plants, just as $K_2CO_3$ is made from the ashes of land-plants, but enough of it could not be had from this source, for it is used in making glass, and soap, and in other large industries.

For these uses it is made from common salt. About half a ton of salt is heated in a furnace with sulphuric acid. This changes the salt into sodium sulphate and sets hydrochloric acid free. Thus:

$$H_2SO_4 + 2NaCl = Na_2SO_4 + 2HCl$$

Then this sulphate, $Na_2SO_4$, also called "salt-cake," is mixed with coal and fragments of limestone, and heated intensely until the whole is melted.

The sulphate is decomposed by the carbon, which seizes its oxygen and flies away as carbon monoxide gas, while sodium sulphide is left behind. Thus:

$$Na_2SO_4 + 4C = Na_2S + 4CO$$

The sodium sulphide, $Na_2S$, and the limestone, $CaCO_3$, then attack and decompose each other, —

$$Na_2S + CaCO_3 = Na_2CO_3 + CaS$$

and a mixture of sodium carbonate and calcium sulphide, $Na_2CO_3$ and $CaS$, called "black-ash," is left.

The carbonate is then dissolved out of the black-ash by water, after which, by evaporating the water, the carbonate comes out as crystals. The sodium carbonate made in this way is known in commerce as *soda-ash.*

There is another carbonate of soda which is made by running carbon dioxide through a solution of $Na_2CO_3$: it is the acid carbonate, $NaHCO_3$. This is "baking-soda."[1] It is much used in "baking-powders," and to furnish the carbon dioxide gas for making "soda-water."

**Sodium Hydroxide. NaHO.** — This substance is so much like potassium hydroxide that either may be used for the same purposes in the laboratory and in the arts. Both are used in soap-making. They combine with the acids which are in fats and oils, and the *salts* which are thus made are called *soaps.* There is this difference, however: the potassium hydroxide makes the soft soaps, while the sodium hydroxide yields the hard soaps.

The "caustic soda," or sodium hydroxide, is made from $Na_2CO_3$ in the same way (how?) that "caustic potash" is made from $K_2CO_3$.

The compounds of sodium are *all* quite freely soluble in water. (Can a precipitate be obtained from a solution of any of them by tartaric acid, as in Ex. 128?)

[1] In the exercises, pp. 39 and 40, the student must have come to the conclusion that "baking-soda" is a sodium carbonate. Having now *gathered more facts,* he finds that his conclusion was true, but that it was not the whole truth. Wholly reliable conclusions can never be reached until "the facts are all in."

**Flame Color.** — The flame of sodium is pure yellow, and any of its compounds decomposed, by heat, will yield this color intensely. Thus:

*Ex. 131.* — Wash the platinum loop and hold it in the Bunsen flame. Is not the flame tinged yellow? Wash the wire, and burn it repeatedly until it does not give the yellow tint, and then use any sodium compound. Note the rich color it imparts.

Look at this colored flame through cobalt-blue glass: is the yellow color to be seen?

There is a little sodium in the air, in the dust of the room, and almost everywhere. The yellow tint of sodium may be given by this small quantity that is ever present. Only a deep, rich yellow can be trusted to show the presence of a sodium salt in quantity.

*Ex. 132.* — Make a mixture of a sodium and a potassium salt and burn the mixture on the platinum loop.

Which color, yellow or violet, can you get with the naked eye?

Which, if you look through cobalt glass?

*Ex. 133.* — Make a strong solution of some of the mixture, and then add a little tartaric acid. Does a precipitate form in the liquid? Is this precipitate made by the potassium or the sodium salt? What is it?

### AMMONIUM (?) $NH_4$.

We have seen that from ammonia, $NH_3$, we may get a class of salts called the ammonium salts. Is there a metal in these salts?

**Some Facts.** — Ammonia, $NH_3$, is very soluble in water (Ex. 56), and this solution will restore reddened litmus to blue, and it also neutralizes acids. This shows

that it is a *base*, and yet it is made up of nothing but nitrogen, hydrogen, and oxygen, — all non-metals, —

$$N H_3 + H_2O = N H_4 H O.$$

Ammonia combines with hydrochloric acid (Ex. 9), to make what is called ammonium chloride. This is a white solid, which has all the properties of a *salt*, resembling the chlorides of sodium and potassium, and yet it is made up of nothing but nitrogen, hydrogen and chlorine, — all non-metals.

Ammonia solution neutralizes nitric acid and forms ammonium nitrate (Ex. 59), and this is as much a *salt* as is any other nitrate, but yet it is made of only nitrogen, hydrogen, and oxygen, — all non-metals. We find no metal in these ammonium compounds.

**Comparison of Formulas.** — Let us now compare these compounds with those which do contain a metal, — say the metal potassium.

Potassium hydrate, $K H O$, and Ammonium hydrate, $N H_4 H O$
Potassium chloride, $K Cl$, and Ammonium chloride, $N H_4 Cl$
Potassium nitrate, $K N O_3$, and Ammonium nitrate, $N H_4 N O_3$

Notice that instead of K in the potassium compounds we find $N H_4$ in the ammonium compounds. The $N H_4$ acts just like the metal potassium in making salts, and this leads us to say that $N H_4$ may be a metal also.

A HYPOTHETICAL METAL. — But we cannot get the $N H_4$ separate, as a metal. These atoms hang together well while in the salts, and in reactions, but part at once if driven out together. In chemical changes the group acts as if it were a metal, and we may *suppose* it to be one. We call it ammonium, and may give it a symbol, Am.

**Its Salts.** — The ammonium salts are numerous. Ammonium chloride, Am Cl, Ammonium carbonate, $Am_2 C O_3$,

and ammonium sulphide, $Am_2S$, may be mentioned. The source of these salts in commerce has been already given, and should now be recalled by the student, p. 85.

THE SULPHIDES. — Only the sulphides need a word more. There is more than one ammonium sulphide, but they are all made by dissolving hydrogen sulphide gas in ammonia water. Large volumes of gas will be absorbed. Thus:

$$H_2S + NH_3 = NH_4HS, \text{ — Ammonium hydrosulphide.}$$

But if ammonia is also present with this $NH_4HS$, they combine. Thus:

$$NH_4HS + NH_3 = (NH_4)_2S, \text{ — Ammonium sulphide.}$$

The solution of these two sulphides is colorless, but if any free sulphur is in it, the liquid becomes yellow. The sulphur unites with the sulphide. Thus:

$$(NH_4)_2S + S = (NH_4)_2S_2, \text{ — Ammonium disulphide.}$$

This last-named sulphide is yellow.

The ammonium sulphide solution is one of the most useful reagents, as the student will soon find. But this reagent, called ammonium sulphide, is not a single compound; it is a mixture of all these. We may study the ammonium compounds further, and compare them with potassium and sodium, as follows:

*Ex. 134.* — Add some tartaric acid solution to a strong solution of any ammonium salt.

Note whether a precipitate comes as with potassium.

*Ex. 135.* — Find out whether an ammonium salt will give any particular flame-color.

*Ex. 136.* — Mix a little solid or liquid ammonium salt, of some kind, with a little KHO, and gently heat it in a test-tube.

Notice the odor; what substance is set free?

*Ex. 137.* — Place a little *solid* ammonium salt in a porcelain dish, and gradually heat it.

Does it melt? What change does occur?

THE FACTS. — Tartaric acid gives a white precipitate in strong solutions of ammonium salts, because the acid tartrate of ammonium is not very soluble. But the ammonium salts generally are very soluble, and on this account they need not be expected to give precipitates. They give no color to the flame. They are decomposed when heated with caustic potash, and then yield ammonia gas, known by its odor. Heat alone drives them into vapor completely, leaving nothing behind.

The change of a solid directly to vapor, without first melting, is called *sublimation.* Camphor is the most familiar example of substances which *sublime* when heated. The ammonium salts, as a class, do this.

## THE POTASSIUM GROUP.

Potassium, sodium, and ammonium (?) are very much alike, and besides these there are three other metals, whose salts are less common and useful, also like these in chemical character. They are lithium, cæsium, and rubidium. These six form a single family of metals, generally called the *Metals of the Alkalies.*

These elements are all soft, light, and silvery white. They lose their luster at once in air, because they have so strong attraction for oxygen that the surface is tarnished with oxide. They cannot, therefore, be kept in air, but must be put up in naphtha, a liquid which has no oxygen in it. They decompose water whenever they touch it, and form bases by the action.

These are the most powerful bases known, — the most caustic and the most ready to neutralize acids. They

form the same classes of salts, and these salts are much alike in properties. These metals are all univalent.

QUERY. — By what experiments would you find out whether a given salt is a potassium, a sodium, or an ammonium compound?

APPLICATION. — Take from the teacher, or a friend who knows what they are, some substances, and see if you can decide whether each is a potassium, or a sodium, or an ammonium compound.

## THE CALCIUM GROUP.

### CALCIUM. *Ca″*.

THE metal calcium is about as hard as gold, and shines with a yellow luster. It tarnishes quickly in air. No use has ever been made of the element itself, but its compounds are not only very useful, but very abundant.

**Occurrence in Nature.** — Calcium carbonate, $Ca'' C O_3$, is one of the largest constituents of the earth. Marble is its purest natural form, but limestone is chiefly the same thing, and the limestone rocks make up a large part of the earth's crust. Calcium carbonate is the starting-point in the manufacture of the useful salts of this metal.

**The Effect of Heat on the Carbonate.** — When a piece of marble is heated intensely, it will not be much changed in looks; its size will be the same, its color somewhat whiter. But it will be found to be more easily crushed, and it is easily shown that a chemical change has occurred. This may be done by treating it with water. Very soon after being wetted, the stone begins to swell and crack and crumble, while volumes of steam arise, and when all is over a fine white powder remains.

The fact is, that heat decomposes the marble, and drives off carbon dioxide. Thus:

$$Ca CO_3 + \text{heat} = Ca O + C O_2$$

and the white mass left behind is the calcium oxide, Ca O. The common name of this oxide is *quick-lime*.

By this simple chemical process large quantities of quick-lime are manufactured. The furnaces are called lime-kilns.

The white powder made by the action of water on quick-lime is called *slaked-lime.* The water actually combines with the oxide. Thus:

$$CaO + H_2O = CaH_2O_2 + \text{heat}$$

We see by the formula that slaked lime is calcium hydroxide. The heat of this chemical action is remarkable.

Among many uses of slaked-lime, we may mention that of making mortar for building purposes. Mortar is made by mixing slaked-lime and sand, and it is used to cement together the bricks or stones in the walls of buildings. Fresh mortar is very weak; it becomes a strong cement after it is laid up in the walls. The chemistry of the change is this:

Calcium hydroxide absorbs the carbon dioxide of the air, when exposed, and is slowly changed by it into *calcium carbonate,* which is hard and stony, and this, filled in with grains of sand, becomes a strong and solid mass which cements the bricks or stones together.

The calcium hydroxide, $CaH_2O_2$, is slightly soluble in water: the solution is known as *lime-water,* and is very useful to the chemist (Exs. 6, 26, 73), and to some extent in medicine.

**Effect of Acids on the Carbonate.** — Even the weaker acids will decompose marble, and in fact most other carbonates. Water and carbon dioxide are produced by this action and the escape of this last in bubbles is the *effervescence* which always occurs. Turn back, now, to the preparation of carbon dioxide (Ex. 73). We wanted the $CO_2$ then, and did not take account of anything else. But we can now see what the complete reaction really was. Thus: —

$$Ca''CO_3 + 2HCl \text{ became } CaCl_2 + H_2O + CO_2$$

and this illustrates the action of other acids on this substance. Carbon dioxide, water, and a salt of calcium are the products.

**Effect of Water on the Carbonate.** — Pure water will not dissolve this carbonate, but water holding $CO_2$ in solution dissolves it quite readily. Let the $CO_2$ escape, or drive it out by heat, and the carbonate reappears.

This action goes on in nature. Rain-water contains $CO_2$, taken from the air, and as this water runs over the lime-rocks it dissolves and carries some of their substance along. But when this solution is exposed to air, the $CO_2$ escapes, and the water then drops the solid carbonate. This occurs in caverns where the water trickles through their roofs. Each drop leaves a little solid carbonate behind when it falls, and an icicle-like mass slowly grows from the roof. This is called a *stalactite*. Another grows up from the floor, this is called a *stalagmite*.

**The Sulphate.** $CaSO_4 + 2H_2O$. — This is known as *gypsum*, — a crystalline mineral found in some abundance, and, when ground to powder, quite useful as a fertilizer. This mineral gives up its water, $2H_2O$, when strongly heated, and crumbles to a fine white powder called "*Plaster of Paris.*" Wet this plaster of Paris, and it combines with water again and hardens into stone. This curious property makes the sulphate useful for making casts.

*Glass* contains calcium silicate, combined with silicates of sodium or potassium (see p. 185), and *bleaching powder* is a mixture of calcium hypochlorite and calcium chloride. How is it prepared? See p. 141, and Ex. 86.

**To Prepare the Insoluble Compounds.** — Among the many salts of calcium there are several that are not soluble in water, especially in water that contains ammonia, and such can be obtained by precipitation.

*Ex. 138.* — Place about 5 cc. of calcium chloride in a test-

tube, add about as much water, and heat the mixture up to boiling-point. Then add slowly a solution of ammonium carbonate as long as it continues to produce the precipitate. The time to stop may be known by letting the solid settle a little, and then notice whether another drop of the ammonium carbonate has any effect. The white precipitate is calcium carbonate. Am Cl was also made, but stays dissolved in the clear liquid. Thus:

$$Ca\,Cl_2 + (N\,H_4)_2\,C\,O_3 = Ca\,C\,O_3 + 2\,N\,H_4\,Cl$$

A very valuable fact is shown in this experiment. Notice that one carbonate precipitates another. We wanted to make the *insoluble* $Ca\,C\,O_3$, and we did it by using the *soluble* $(N\,H_4)_2\,C\,O_3$. The fact is, that a soluble salt precipitates an insoluble salt of the same class as itself. If we want an insoluble *hydrate* we will use some soluble *hydrate* to make it. Or, if we want an insoluble sulphide we will use some soluble sulphide to produce it. There are a few exceptions to this rule, but it is so generally true as to be an excellent guide.

For example, we wish to make the insoluble calcium sulphate. Let us select some soluble sulphate to do it with. It may be potassium sulphate, $K_2\,S\,O_4$, or hydrogen sulphate, $H_2\,S\,O_4$, or some other.

*Ex. 139.* — Place about 5 cc. of water in a test-tube, and add about as much solution of *any* soluble compound of calcium, — such as the chloride or nitrate, — and then add, little by little, a solution of $K_2\,S\,O_4$.

Repeat the operation, using $H_2\,S\,O_4$.

The precipitate is calcium sulphate in both cases.

Can you write the reactions?

But the term insoluble is not applied to a substance, because the fluid dissolves absolutely none of it; we call a thing insoluble when a fluid dissolves only *very little.*

Everything is soluble in some degree, but in many cases the quantity, which will dissolve in the amount of fluid used, is too small to be taken account of. Such are called insoluble substances.

Now $CaSO_4$ is *somewhat* soluble in water. In fact, 400 cc. of water will dissolve about 1 g. of it, and if you have, say 10 cc. of fluid in your test-tube (Ex. 139), there must be about $\frac{1}{40}$ of a gram of the sulphate in solution; the rest of the sulphate is the precipitate. A precipitate will always be seen whenever the fluid present cannot dissolve all of the substance which the reagent makes, but not otherwise.

**To Prepare Soluble Compounds.** — If, in any case, the new compound made by a chemical reagent is soluble, it may be obtained by evaporating the clear liquid, in which it is dissolved, until crystals will form, or to dryness if necessary. For practice, the student should prepare some calcium nitrate, and calcium acetate, from the pure carbonate, or from marble.

### THE CALCIUM GROUP.

Two other metals, barium, Ba″, and strontium, Sr″, are very much like calcium, and these three form the calcium group of metals, or, as it is also called, the group of the *alkaline earths.*

In general, these three metals combine with the same substances and in the same proportions. Their attraction for oxygen makes them tarnish quickly in air, and enables them to decompose water when they touch it. The hydroxides thus formed are strongly alkaline. In these chemical actions barium is more energetic than strontium, and strontium more than calcium. Now this is also the order of their atomic weights, if we start with the largest. Thus:

Ba = 137 Sr = 87.5 Ca = 40

This is another illustration of the curious fact, noticed among groups of the non-metals, that the properties of elements seem to depend on their atomic weights.

The great resemblance of barium and strontium to calcium will be seen if we make and compare some salts of these three metals. Some differences will also be discovered.

*Ex. 140.* — Make the carbonates of barium and strontium just as that of calcium was made in Ex. 138, and notice how much alike these three carbonates appear.

*Ex. 141.* — Arrange three test-tubes, each with 5 cc. of water, and add to one 5 cc. of *strong* solution of $BaCl_2$, to the second, as much strong solution of $SrCl_2$, and to the third, as much strong solution of $CaCl_2$. Then add to each a little solution of calcium sulphate.

Notice a precipitate *at once* in the Ba tube only. Heat the other two to boiling, and notice then a precipitate in the Sr tube, but none in the other.

The barium sulphate forms at once in the cold.

The strontium sulphate at once only when heated.

The calcium sulphate does not form at all.

This experiment will help us to decide, in any case, which one of these three metals, if either, is present in a given solution. But why is not the calcium sulphate precipitated here as it was in experiment 139?

**Flame Colors.** *Ex. 142.* — A volatile compound of calcium will tinge the flame yellow-red; of strontium, brilliant crimson; of barium, green. Try these compounds, and observe the flames with the naked eye, and also through cobalt glass. Mark well the difference between these and the flame colors of potassium and sodium.

## METALS OF THE ZINC GROUP.

**Magnesium. Mg″.** — The metal magnesium is found as a carbonate and as a silicate, these two being its most important compounds in the rocks. Its sulphate, known as Epsom salt, is its most important compound. It is found in solution, and in some mineral springs it is a valuable constituent. What is commonly called magnesia is the oxide of this metal, $Mg''O$. It is used in medicine. The Epsom salt, $MgSO_4 + 7H_2O$, is still more valuable as a medicine; but on the whole magnesium and its compounds do not rank high among useful substances.

The chemical actions of magnesium compounds are a little like those of the calcium group. Its carbonate is insoluble in water.

*Ex. 143.* — Precipitate $Mg''CO_2$ from its chloride just as was done for $CaCO_3$ in Ex. 138.

*Ex. 144.* — Put 5 cc. ammonium chloride in a test-tube, and to this add 5 cc. of magnesium chloride. Then add the ammonium carbonate. Does a precipitate form? If not it shows that magnesium carbonate is soluble in ammonium chloride. This is a fact. In this respect this carbonate *differs* from those of the calcium group. In general the compounds of this metal are more soluble than those of that group.

### ZINC. *Zn″*.

The metal zinc is found in several minerals which occur in considerable quantities in the rocks, such as "calamine," which is the carbonate, "blende," which is the sulphide, and "zincite", the oxide. These are its most valuable *ores*.

Zinc is a useful metal, and it is manufactured on a large scale from these native compounds.

**Manufacture of Zinc.** — Suppose the metal is to be taken out of the oxide, Zn O. The problem is how to get rid of the oxygen. Turn back to Ex. 72, and you recall the strong attraction of carbon for oxygen, and how, on this account, carbon *reduced* Cu O to metallic copper. Now this is the fact applied to get the zinc out of Zn O on a large scale. The oxide and charcoal are put into vessels of fire-clay and heated in a furnace. What reaction occurs? It may be written:

$$Zn\,O + C = C\,O + Zn.$$

The zinc, in the form of vapor, is led out into cold vessels, where it is condensed.

But suppose the zinc is to be taken out of one of the other ores, — say the sulphide, Zn S. The problem, then, is to get rid of the sulphur. The problem is solved by first changing the sulphide into the oxide, and then reducing the oxide by carbon as before. And to change the sulphide to oxide it is simply heated in the air, when oxygen of the air takes the place of sulphur in the ore. Thus:

$$Zn\,S + 3\,O = Zn\,O + S\,O_2$$

This heating of an ore in the air is called *roasting.* The object of it is to change the ore to an oxide. It is in many cases, as in this one, the first step in the process of getting a metal from its ores.

**Uses of Zinc.** — The uses of zinc are quite numerous. When heated to a temperature above that of boiling water (125° – 150°) it can be rolled out into thin sheets. This sheet-zinc is in familiar use. Zinc is brittle, and cannot be rolled out at temperatures much higher or lower than those given above. Zinc is also used as a covering for

sheet-iron; for this purpose sheets of iron are simply dipped in melted zinc; it is then called *galvanized iron.* Zinc and copper melted together form *brass.* Zinc, copper, and nickel, melted together, form *German silver.* Substances like these, which consist of two or more metals together, are called *alloys.* These and other alloys of zinc are useful.

**Compounds of Zinc.** — Of the compounds of zinc we may mention the following: zinc oxide, which is used as a paint under the name of *zinc white,* and is made for this purpose by heating the carbonate. Zinc chloride, used by tinners for cleaning the surface of metals for soldering, and also as a disinfectant. Zinc sulphate, known as *white vitriol,* a poisonous substance, but used in small quantities in medicine.

**Preparation of Insoluble Compounds.** — Among its compounds which are insoluble in water are the carbonate, the sulphide, and the hydroxide. To make these we may start with zinc chloride.

*Ex. 145.* — Make a solution of zinc chloride for use, by putting bits of zinc in a bottle and pouring a few cubic centimeters of $HCl$ on them. There should be zinc left in the bottle when the action is over, for then the zinc chloride will be free from acid.

*Ex. 146.* — Add a few drops of this solution of $ZnCl_2$ to 10 cc. of water, and then add $(NH_4)_2CO_3$.

Describe the result and write the reaction.

*Ex. 147.* To 10 cc. of water, with a few drops of $ZnCl_2$ add $Am_2S$.[1] What is the *white* precipitate which falls?

[1] This ammonium sulphide, and the hydrogen sulphide for the next experiment, are made by passing $H_2S$ through dilute ammonia (half water) for one and water for the other. Use the same apparatus as in Ex. 107.

Write the reaction. What soluble compound is also made in this reaction?

*Ex. 148.* To a few cubic centimeters of $Zn\,Cl_2$ add a few drops of hydrochloric acid, and then add a solution of hydrogen sulphide, $H_2S$. Do you get a precipitate? If not, it shows that $Zn\,S$ is soluble in this liquid.

We make the sulphide whether we use $Am_2S$ or $H_2S$, but the first is an alkaline substance, while the second is an acid, and we find that the sulphide is not soluble in the *alkaline* liquid, Ex. 147, while it is soluble in the *acid* liquid, Ex. 148.

*Ex. 149.* — To a few cubic centimeters of the $Zn\,Cl_2$ with an equal bulk of water add $N\,H_4\,H\,O$ little by little, until after shaking it, and then blowing the air out of the tube, the strong odor of ammonia remains.

Describe the two changes which occurred.

The explanation is this: Ammonium hydroxide changes the zinc chloride to zinc hydroxide, which takes the form of a white precipitate because it is *insoluble in water.* But it is *soluble in ammonium hydroxide,* and so just as soon as there was added more than enough to make the precipitate, the excess began to dissolve that which had been made. Whenever, as in this case, a precipitate dissolves in the reagent which makes it, it is said to be *soluble in excess.*

**The Zinc Group.** — Magnesium, zinc, and cadmium are the members of this group. These elements are much alike. They are all bluish-white metals. They have many of the same properties, but in different degrees.

For example, cadmium melts at a comparatively low temperature, zinc at a higher; 423° C., and magnesium at one still higher. Cadmium becomes a vapor at a low red heat, zinc if heated but little hotter, and magnesium at a bright red heat. They take fire in the air, cadmium

burns less freely, zinc with a fine blue flame, and magnesium with a dazzling whiteness. They decompose dilute acids, cadmium rather slowly, zinc more freely, and magnesium very promptly.

This order of their properties is also the order of their atomic weights, beginning with the largest. Thus:

| Cd | Zn | Mg |
|---|---|---|
| 112 | 65 | 24 |

SUGGESTION. — Compare, by experiment, the action of $H_2S$, and of $Am_2S$, on some soluble compounds of zinc, cadmium, and magnesium, and note the different results.

Then take from the teacher or a friend, who knows what they are, some specimens, and see if you can decide whether each is a compound of either of these metals.

## THE IRON GROUP.

### MANGANESE. *Mn″*.

As a metal, manganese need not detain us, since it is rarely used in chemistry or the arts. In the form of the "black oxide," $MnO_2$, it is found in many parts of the earth. This is its chief ore, although it occurs in other minerals and rocks in some abundance. It unites with oxygen in several proportions, reminding us, in this respect, of nitrogen: there are five oxides of each.

**Salts of Manganese.** — Two of these — those which contain least oxygen — act with *acids* to produce salts: they are basic, while one of them — that which contains most oxygen — acts with *bases* to produce salts: it is acid. The other two do not produce salts at all: they are neutral. The student should think of the bearing of this fact on our definition of metal, p. 192.

From the basic oxide, MnO, we may get manganous chloride and manganous sulphate, which are the salts most common in the laboratory, while from the acid oxide, $Mn_2O_7$, we may get the potassium permanganate, which is a beautiful salt, much used in the laboratory and out of it. Its solution in water is intensely purple, but this fine color is quickly lost in presence of anything which has an affinity for oxygen. It parts with its oxygen so readily to other things, that it is a most powerful *oxidizing agent*. In the laboratory, it is used for this purpose, and out of it too; for, since bad odors and putrid organic matter are destroyed by oxygen, the permanganate is a valuable disinfectant.

Preparation of Some Insoluble Salts. — Among the insoluble salts of manganese we may mention the mangan-

ous carbonate, sulphide, and hydrate. A study of these will reveal some of the most characteristic reactions of this metal.

*Ex. 150.* — To make the carbonate, proceed exactly as in Ex. 146 with zinc, using a solution of manganous sulphate instead of the zinc compound.

What is the color of manganous carbonate?

What is the color after some time in the air?

To answer this last question, it is well to filter the liquid, and leave the precipitate on the filter. The change you will discover is caused by the oxygen of air, which changes the manganous carbonate to the manganic hydrate. This change does not occur in the case of zinc.

*Ex. 151.* — To make the sulphide, proceed just as in Ex. 147 with zinc, using manganous sulphate, and $Am_2S$.

What is the color of the manganous sulphide?

Write the reaction which took place.

What soluble compound is made at the same time?

*Ex. 152.* — Try to make the sulphide by using $H_2S$. Proceed exactly as in Ex. 148 with zinc.

What is the result, and why?

*Ex. 153.* — To make the hydrate, $Mn H_2O_2$, proceed just as in Ex. 149 with zinc, using $Mn SO_4$.

What is the color of this manganous hydrate?

Is it, like the zinc hydrate, soluble in excess?

What change happens if it is left in the air?

The pure manganous hydrate is white, but it is changed by oxygen into brown manganic hydrate.

*Ex. 154.* — To a few cubic centimeters of the manganous solution, add as much ammonium chloride, and then add ammonia as before.

Do you get the precipitate? If not, it shows that manganous hydrate is soluble in ammonium chloride. You will get none of it, if enough of this chloride is present.

PREPARATION OF POTASSIUM MANGANATE. — The manganate is soluble in water, and of course it cannot be precipitated. But there is another way of making salts, and that is *by fusion*, or, as it is often called, the "dry way."

Let us mix a little "black oxide" of manganese with an equal weight of solid potassium hydroxide and half as much potassium chlorate. This mixture should be strongly heated. It may be done in a thin iron spoon, or the bottom of a broken porcelain dish. When well fused, the mass will turn green, and this green substance is potassium manganate, $K_2MnO_4$. Let it cool and then put it into water, and you shall find that the manganate dissolves, yielding a fine green solution.

PREPARATION OF PERMANGANATE. — If now you boil some of this green solution, you shall see its color change to a rich purple-red. The fact is, that the green manganate is easily *oxidized*, and is thus changed into the red permanganate, $K_2Mn_2O_8$. This change takes place on boiling its solution. The curious change of color, from green to red, has given to the manganate the name "chameleon mineral." The permanganate holds its oxygen very loosely, and will therefore oxidize other bodies readily. By this action it decomposes organic matter. When added to water containing organic matter, it becomes colorless or brown by giving up its oxygen.

**Nickel.** Ni''. — Nickel is a white, hard metal, which is much like iron, but does not rust as easily, and on this account it is used as a plating on the surface of other metals, such as steel, copper, and brass.

Nickel salts are generally green, but the description of them need not detain us.

**Cobalt. Co″.** — Cobalt is more rare than nickel. It also is a hard, white metal, with a reddish tinge however, and in other qualities it is much like nickel and iron. The salts of cobalt are highly colored, and the color depends on whether they contain water. Thus cobaltous sulphate, when it contains water, $CoSO_4$, $7H_2O$, is a beautiful red salt, but once drive its water off by heat, and the sulphate, $CoSO_4$, is blue. The silicate is used to color glass; so is the oxide. For this purpose no substance yields a richer blue. The blue glass used in viewing colored flames is an example.

The student should prepare some nickel and cobalt salts and study them by making, with them, the same experiments as have been already made with salts of zinc and manganese.

### IRON. $Fe''$ or $(Fe_2)^{VI}$.

While manganese, nickel, and cobalt are quite rare and little used, iron is the most abundant and the most useful of metals.

**Occurrence in Nature.** — Iron is found in the rocks, in the soil, in plants, and in the bodies of animals. There is no metal more widely diffused than iron.

Native iron is sometimes found in the earth, and "meteoric stones," which come from space outside the earth, are little else than iron, with also some native nickel and cobalt.

The sulphide, $FeS_2$, "pyrites" is found almost everywhere, sometimes in yellow scales, sometimes in perfect cubes, which are also yellow. It is so often mistaken by the ignorant for a precious metal, that it is called "fool's gold."

The chief *ores* of iron are two oxides and a carbonate. The richest of all is the "Magnetic oxide," $Fe_3O_4$, which is so called because it is able to attract a magnet. It is very

abundant in this country. There are mountains of it in Missouri, and vast beds of richest quality in New York. It is black.

Hæmatite is the other oxide, $Fe_2O_3$. It is red or brown, and sometimes, in the form of beautiful crystals, its color is dark steel-gray.

The carbonate, $FeCO_3$, is still less rich in iron, and it is generally mixed with clay and other earthy matters, which make it poorer still. Such a carbonate is called "clay-iron stone."

These ores are rarely pure; they are largely mixed with sulphides and earthy matter, and the red oxide contains water; it is $Fe_2O_3, 3H_2O$. All these facts have to be taken into the account in the process of extracting the iron.

Roasting the Ores. — In working other ores than the best oxides, they are first roasted, p. 213. In this process water is driven off, sulphur is burned into sulphurous oxide, carbon dioxide is set free from the carbonate, and the iron of the whole is changed into oxide, which remains mixed with the earthy matters of the ore.

Reducing the Oxide. — The next thing to be done is to decompose the oxide and liberate the iron. The power of hot carbon to reduce an oxide, Ex. 72, is here applied.

The process is carried out in a *blast-furnace*, — a furnace in which the fire is driven by a blast of air. Such a furnace is pictured in Fig. 64. It may be 50 or 100 feet high, and at its widest place inside it may be 12 or 18 feet in diameter. The furnace is fed from the top, and kept full of fuel, crushed ore, and fragments of limestone, thrown in together.

The fire is started at the bottom, and is urged to the greatest intensity by a blast of air driven in by the power of a steam-engine.

In this intense heat the earthy parts of the ore, and the

limestone, melt together into a glassy substance, called the "slag." The carbon of the fuel seizes the oxygen of the ore and sets the iron free. Kept melted by the intense heat, the iron runs to the bottom of the furnace into a sort of chamber made to receive it, while the lighter slag floats on the surface of the melted metal.



Fig. 64.

In front of the furnace is a large level bed of sand, with one main furrow through its center, and many branch furrows on either side, as shown in the cut.

At intervals of about twelve hours the furnace is opened at the bottom for the metal to run out. It then flows down into the furrows of the sand-bed, where it is allowed to grow cold.

The iron thus made is called *cast-iron*, and, in the form of the short bars made in the sand, it is called *pig-iron*.

But besides iron this cast-iron contains carbon, silicon, sulphur, and phosphorus, in small quantities. These impurities make the metal weak, brittle, and fusible. Nevertheless, it is used for the manufacture of a great many articles where great strength is not required.

In fact, pure iron is not found at all in commerce. The very purest contains a little carbon.

**The Three Forms of Iron.** — There are three forms of iron, known as cast-iron, wrought-iron, and steel.

WROUGHT-IRON. — This is the purest form of commercial iron; it is the toughest, strongest, and most malleable. It is made from cast-iron by robbing it of its impurities, and this is done by bringing oxygen into contact with all parts of the melted metal. The furnace in which this is done is called a *reverberatory* furnace, and by help of the cut, Fig. 65, the operation may be understood.

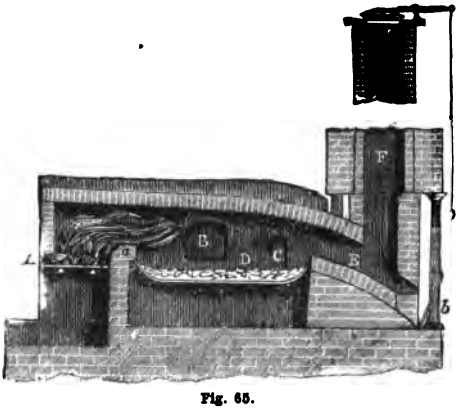


Fig. 65.

The cast-iron is placed upon a large hearth, *D*; the fire is built in a separate part of the chamber. Flame and hot gases from the fire strike against the arched roof of the furnace, and the intense heat is thrown from the roof down upon the cast-iron. In a little time the iron begins to soften, and at length it becomes a pasty mass of half-melted metal. Then the furnace-man unstops a hole, *B*, thrusts a paddle through it into the pasty mass, and working this paddle about (puddling) he thoroughly stirs the metal, so that all parts of it are slowly brought in contact with the hot air at the surface.

The hot oxygen of the air seizes the impurities, one after another. Some of the new compounds form a liquid slag, which is drawn off out of the furnace at *b*, while others are gases which pass up and out of the high chimney.

The result is, that the iron is soon left almost free from its impurities. A large part of its carbon has been burned away, and has gone off as carbon dioxide. A large part of the other impurities are also burned, but some of the slag, thus made, is still mixed with the purified iron.

The furnace-man then lifts out a ball of this pasty iron, weighing perhaps sixty pounds, or even more, and puts it under the heavy blows of an immense hammer, or sometimes under the tremendous pressure of the squeezer. In this way the impurities which are still mixed with the iron are pounded or squeezed out, so that the iron is left more pure and compact.

The mass of iron is then forced through grooves between two strong rollers. The very great pressure of these rollers lengthens the mass out into a slender rod or bar.

For the best quality of *bar-iron*, as this form is often called, the bars, made by the first rolling, are cut into short lengths and bound together in bundles, to be heated over again and rolled into bars a second time. By repeating these efforts the purest and strongest iron to be found in commerce is obtained. It is *wrought-iron*, also called *malleable iron*.

STEEL. — Neither *cast-iron* nor *wrought-iron* has the qualities which render iron useful in all the various ways in which this metal is employed; a third form, which is better than either of these for many uses, is *steel*.

The difference between steel and the other two kinds of iron is best shown by the action of fire and water.

Let wrought-iron be strongly heated and then suddenly cooled by a plunge into cold water. Very little if any change will be produced.

Let steel be treated in the same way, and it will be made almost as hard as diamond, and so brittle that it will snap before it will bend. Let this brittle steel be heated again

to a point below red-heat and then cooled, and it is softened somewhat, and, what is more remarkable, it is made so elastic that it will bend rather than break, and spring back again when released from the force.

Cast-iron may be hardened as much as steel may be, but it cannot be made elastic.

With this difference in properties we find a difference in composition, for while each is composed of iron and carbon, it is found that *cast-iron* contains the largest proportion of carbon, *wrought-iron* the smallest, and *steel* a proportion between the other two.

It would seem, then, that steel could be made in two ways:

By *taking away* some carbon from cast-iron.

By *adding* some carbon to wrought-iron. And, in fact, the steel so largely used in the arts, is made in both these ways. Thus:

1. From two to six tons of cast-iron is melted and then run into a large globe-shaped vessel made of a substance which will not melt at the highest heat to be attained. In the bottom of this vessel there are many holes, and a strong blast of air is driven through them. The air bubbles up through the melted metal, and a most furious burning begins. The oxygen attacks the silicon, and the sulphur, and the carbon, and some of the iron also, and burns them into compounds with itself.

In this way the iron is partly purified, but the chemical action goes too far, and removes so much carbon that too little is left. By adding some cast-iron to the purified mass, enough carbon is given back to change the whole into steel.

The globe-shaped vessel (*converter*) is then turned on its pivots, and the melted steel is run out into a ladle, and then poured into moulds.

Less than half an hour is time enough to change these tons of cast-iron into steel. The process is called the *Bessemer* process.

2. In the other way of making steel, bars of wrought-iron are packed in charcoal, and the two are shut up together in air-tight boxes. They are then made red-hot, and are kept so for several days.

During this heating, carbon finds its way into the solid iron and changes the whole mass to steel. This product is known as *blistered steel*, because the bars, on coming out of their hot bed, are found to have a great many bubbles or blisters on their surface. The steel is then melted and run into moulds. This method of making steel is called *cementation*.

**Compounds of Iron.** — Some facts about the chemical action of iron have been noticed already. The production of ferrous sulphate in Ex. 120, and of ferrous chloride in Ex. 94, showed the action of iron on acids, — the most direct way of making some salts of this metal.

Experiments 96 and 97 are still more instructive. They should now be studied over again, because they *prove* that there are two distinct chlorides of iron, — the ferrous chloride, $FeCl_2$, and the ferric chloride, $Fe_2Cl_6$. And this illustrates a most important fact, for iron forms not only two chlorides, but also two sulphates, two nitrates, and, in fact, two large classes of salts, — the ferr*ous* salts, such as ferrous chloride, and the ferr*ic* salts, such as ferric chloride.

Compare the Ferrous and Ferric Compounds. — These two classes of iron compounds are unlike in appearance, and quite different in their chemical behavior. These differences are most distinctly shown by experiments. In the first place

*Ex. 155.* — Make some ferrous and some ferric chloride,

to be used in the work which follows. To do this use clippings of small iron wire, or small nails, and add them to hydrochloric acid and aqua regia *until these liquids will dissolve no more.* For directions, see Ex. 96 and Ex. 97. The hydrochloric acid yields solution of ferrous chloride. The aqua regia yields solution of ferric chloride.

Note the colors of these solutions.

Ferrous salts are generally light green; while ferric salts are generally reddish yellow or brown.

*Ex. 156.* — Prepare two test-tubes, each with about 10 cc. of water, and add to one, 1 cc. of the ferrous chloride and to the other, 1 cc. of the ferric chloride. Now add ammonium hydroxide in drops to the first, and notice the precipitate which forms. Afterward treat the second in the same way. The ferrous chloride yields ferrous hydroxide, $Fe(HO)_2$. The ferric chloride yields ferric hydroxide, $Fe_2(HO)_6$. The color of the first is light green, of the second brownish red.

Now filter the ferrous precipitate out, and leave it on the paper in the air. You will find it turning red like the other hydroxide. In fact, the *moist* ferrous hydroxide actually changes into the ferric hydroxide. Thus:

$$2\,Fe(HO)_2 + H_2O + O = Fe_2(HO)_6.$$

So it is also with other iron compounds of the *ous* class: the air oxidizes them into the *ic* form. This illustrates the rusting of iron in moist air: this same reddish hydroxide is made. We then call it iron-rust.

But the *ous* compounds become *ic* compounds by the action of other oxidizing agents than air, — such as nitric acid. Thus:

*Ex. 157.* — Put about 1 cc. of ferrous chloride with about 5 cc. of water in a tube, add 4 or 5 *drops* of strong $HNO_3$, and then boil it gently for a minute.

Notice a change in the color of the solution.

What change in the substance does this color suggest?

Now add ammonium hydroxide, and observe the precipitate. Is it green or red, — ferrous or ferric hydroxide?

What change does this prove that the $H\dot{N}O_3$ made?

The nitric acid will, in the same way, change other *ous* iron compounds into the *ic* form. It is said to *oxidize* them, and yet it does not change them into oxides. It changed the $FeCl_2$ into the $Fe_2Cl_6$, and it would change $FeSO_4$ into $Fe_2(SO_4)_3$. Whatever changes any compound from a lower to a higher form of combination is called an *oxidizing agent.*

*Ex. 158.* — Put two or three drops of ferrous chloride with about 5 cc. of water, and then add *drops* of a solution of potassium ferrocyanide.

In another tube make the same experiment with ferric chloride. Note the different results carefully.

A beautiful deep blue precipitate is always made by potassium ferrocyanide in ferric salts: it is "Prussian blue." The same reagent in ferrous salts yields a pale blue precipitate, and this difference is so marked, that one can, in this way, tell whether a given iron solution holds a ferrous or a ferric salt.

COMPARE SOME COMPOUNDS OF IRON WITH THOSE OF ZINC AND MANGANESE. *Ex. 159.* — Put a few drops of ferric chloride in water, and add ammonium hydroxide "in excess," to see whether ferric hydroxide is soluble in excess, like that of zinc, Ex. 149.

*Ex. 160.* — Put a few drops of ferric chloride into a little water, add considerable ammonium chloride, and then the ammonium hydroxide, to see whether the ferric hydroxide is soluble in Am Cl, like that of manganese, Ex. 154.

*Ex. 161.* — To 10 cc. of water with 1 cc. of ferric chloride add a few drops of ammonium sulphide, $Am_2S$.

What is the name and the color of this precipitate? Compare it with the sulphides of zinc and manganese.

*Ex. 162.* — To 1 cc. of ferric chloride add a drop or two of hydrochloric acid, and then add a solution of hydrogen sulphide. Does the color of this result agree with that of any of the iron compounds seen before?

Ferrous sulphide, Fe S, is a black substance which is not soluble in alkaline liquids, but is soluble in acids. For this reason it is made by $Am_2S$, but not by $H_2S$. In this respect it is like the sulphides of zinc and manganese. The *whiteness* of the liquid in Ex. 162 is due to *sulphur*. The $H_2S$ is decomposed by the ferric chloride, and its sulphur is set free.

**Chromium. Cr.** — Chromium may be made by heating its chloride with potassium, it is then a dark gray powder, which in air takes fire before it reaches a red heat. But if the chloride is heated with sodium, instead of potassium, the chromium is set free as hard and shining crystals. While, if the metal is made from its oxide, by heating it with charcoal, it has a steel-gray luster, is hard enough to cut glass, and combines with oxygen slowly when heated in the air. It is a curious fact that, if we make the same substance in different ways, it sometimes comes out with different properties, as in this case of chromium.

In its chemical actions, chromium is sometimes a metal and sometimes not. With acids it forms salts, like other metals, such as chromium sulphate, $Cr_2(SO_4)_3$. But with bases it forms salts, like the non-metals, such as potassium chromate, $K_2CrO_4$. Is it nature, or only the chemist, who divides the elements into metals and non-metals?

The chief ore of this metal is *chromite*, a compound of iron chromium and oxygen. It is not the metal itself which is obtained from this chromic iron, but its more useful compounds.

Let the chromite be heated with a mixture of potassium carbonate and potassium nitrate, and the potassium will take the place of the iron. In this way the ore, iron chromite, is changed into potassium chromate, $K_2CrO_4$. This new chromate can then be washed out with water. The solution is very yellow, and when evaporated yields a highly-colored yellow salt.

Now let a solution of this $K_2CrO_4$ be treated with nitric acid. Its color will change from yellow to red (try it), and its substance will change from potassium chromate to potassium dichromate, $K_2Cr_2O_7$. By evaporating the solution, fine orange-red crystals may be obtained. (Do this.)

**Compounds of Chromium.** — The chromates generally are highly colored, and some of the insoluble ones are used as paints. The lead chromate, or "chrome-yellow" as it is known in commerce, is a good example.

*Ex. 163.* — Make a solution of either potassium dichromate or chromate, in water, and add little by little some solution of lead acetate. The bright yellow precipitate is the pigment, chrome-yellow, $PbCrO_4$.

Salts, such as chromium sulphate, in which the chromium is basic, are sometimes violet-colored, sometimes green.

Chrome alum is a quite common substance made of two sulphates, the chromium and potassium sulphates. It has a fine violet color. From this compound we can easily get chromium hydroxide, which is green. Thus:

*Ex. 164.* — To a solution of a little chrome alum in water, add, little by little, some ammonium hydroxide. Notice the bluish-green precipitate of $Cr_2(HO)_6$.

Is this, like the zinc hydroxide, soluble in excess?

Is it, like the manganese hydroxide, soluble in Am Cl?

Will it, like the ferrous hydroxide, turn red when exposed to air?

Can you get the same precipitate by K H O instead of $NH_4HO$?

And if so, is it soluble in excess of K H O?

*Ex. 165.* — To a solution of a little chrome alum in water add $Am_2S$, little by little, and compare this precipitate with that in the last experiment. It is the *same substance.*

There is a chromium sulphide, $Cr_2S_3$, but it cannot be made in the presence of water, like the sulphides of manganese, zinc, and iron; the hydroxide will come instead.

### THE IRON GROUP.

The five metals — iron, manganese, nickel, cobalt, and chromium — are closely related, and, together, form the Iron Group.

The metals in this group are much alike in color, luster, and hardness, and much alike in chemical behavior. For example, they all unite with oxygen, and in more proportions than one, making oxides which are sometimes basic and sometimes acid.

But their affinity for oxygen is not equally strong, for while chromium and manganese rust quickly when exposed to moist air, iron does so much more slowly, while nickel and cobalt will keep their luster unless heated. It is interesting to note that this order of the metals, in their behavior with oxygen, is also the order of their combining weights. Thus:

| Cr | Mn | Fe | Ni | Co |
|---|---|---|---|---|
| 52 | 55 | 56 | 58 | 59 |

Another fact is curious: these atomic weights are nearly alike. In other groups we have not found it so. Look, for example, at the potassium and the calcium groups.

SUGGESTIONS. — If the student has made the experiment with the ammonium sulphide, $Am_2S$, he has seen how very different are the sulphides of zinc, manganese, iron, and chromium. And if he has used ammonium hydroxide, $NH_4HO$, he has also seen how very different are the hydroxides of these metals. Now, by the experiment with these two reagents, he ought to be able to tell with considerable certainty whether any substance which is given him is a compound of one or another of these metals.

Let him try to do this by taking specimens from the teacher or a friend who knows what they are. Nickel and cobalt may also be included in the list, if thought best.

If the substance given will not dissolve in water, perhaps you can change it into another compound, of the same metal, which will. Ferrous sulphide is not soluble in water, but hydrochloric acid will change it into ferrous chloride, which is. Now, suppose you do not know that the specimen is ferrous sulphide, and yet you wish to find out if it is a compound of iron or of some other metal. You can treat it with a little hydrochloric acid and get a solution of the chloride, and then you can use the $Am_2S$ and the $NH_4HO$.

If hydrochloric acid will not answer, you can use nitric acid, or even aqua regia, to get the substance into solution. But in all cases, *just as little of these acids as possible* should be used.

## ALUMINUM. *Al*″.

ALUMINUM is a beautiful blue metal, with a luster like silver. It can be hammered into thin sheets, or drawn into fine wire, or cast into any desired form, like iron. It is one of the light metals; it is only 2.56 times heavier than water. It does not easily tarnish in air, and it melts only at a high temperature.

At the same time that it has all these valuable properties, aluminum is one of the most abundant metals. It is found combined with silica in clay, and in all slate rocks, which are little more than hardened clay. In fact, about one-twelfth the weight of the solid parts of the earth is aluminum.

But there is no cheap way to get this metal from its ores, and it is, therefore, too costly to be used in the arts, in place of silver and iron, for many purposes, as it would otherwise be. Its use is limited to ornamental work, and to small articles of apparatus where strength with lightness are required.

**Compounds of Aluminum.** — *Alum* is the most useful compound of this metal. Alum is really a compound of two sulphates and water, for it is potassium and aluminum sulphate with much "water of crystallization." Its formula is $K_2Al_2(SO_4)_4 + 24H_2O$. This is the common alum, although "ammonium alum" is also much used instead; it has ammonium in place of potassium.

*Aluminum oxide*, or alumina as it is usually called, is found in the rocks in beautiful crystals, having different colors and known by different names. The *topaz* and the *emerald* are examples; the first is yellow and the second green, but both are alumina. So are the oriental *amethyst*,

which is purple, the *sapphire*, which is blue, and the *ruby*, which is red.

Some compounds of aluminum are largely used in operations of calico printing and dyeing. This is true of the sulphate and the hydroxide.

The following experiments will reveal some of the chemical characters of the aluminum compounds.

*Ex. 166.* — Make a dilute solution of alum, and add, little by little, $NH_4HO$. The white gelatine-like precipitate is aluminum hydroxide, $Al_2(HO)_6$.

Is this $Al_2(HO)_6$ soluble in excess?

Can you get it by means of $KHO$ instead of $NH_4HO$?

And if so, is it soluble in excess of $KHO$?

*Ex. 167.* — Make a dilute solution of alum as before, and color it just distinctly red with a solution of cochineal. Then add $NH_4HO$. The aluminum hydroxide now combines with the coloring-matter of the cochineal. It is red, while the liquid is left colorless.

The color cannot be washed out of the hydroxide. Now, it is this fact which makes the aluminum salt useful in dyeing. For if cloth is first soaked in a solution of alum and coloring-matter, and then plunged into a solution of ammonia, the same reaction will take place *in the fibers of the fabric*, from which the color cannot afterward be washed out by water.

*Ex. 168.* — To a dilute solution of alum add a few drops of $Am_2S$, and compare the precipitate with that in Ex. 167. It is the same substance.

There is a sulphide of aluminum, $Al_2S_3$, but water decomposes it at once, changing it to the hydroxide. On this account the sulphide cannot be made in solutions. In this respect aluminum is like chromium.

## THE ANTIMONY GROUP.

**Antimony.** **Sb'''.** — This metal is now and then found native, but it is oftener found in combination with sulphur or oxygen. The chief ore is the antimony sulphide, $Sb_2S_3$.

The metal has a silvery appearance, quite brilliant, but, unlike silver, it is very brittle. Its alloys are more important than itself. We only need to mention type-metal, which is far the most important of all alloys, since the art of printing depends upon it. Antimony, tin, and lead are melted together to make this alloy, from which all types for printing are cast.

There are three oxides of antimony, and these are all inclined to form acids rather than bases. This shows that antimony is more like the non-metals than any other metal which has yet been described. Another evidence of this is the fact that antimony combines with hydrogen, just as do nitrogen and phosphorus. The compound is called stibine, and its formula is $H_3Sb$. Stibine may be made and burned just like arsine, Ex. 123, and it then yields a stain of antimony on porcelain by Marsh's test, p. 181.

**Bismuth.** **Bi'''.** — This metal is crystalline, brittle, and bright, with a reddish color. Bismuth, like antimony, forms an alloy with lead and tin. This alloy is called *fusible metal*, and it merits this title because it melts at a temperature of only 94° C. The separate metals must be heated much higher: tin melts at 228°, bismuth at 267°, and lead at 325°, but this alloy at 94°. Such is very generally the case with alloys: they melt more easily than any one of the metals of which they are made.

Fusible metal, like type-metal, expands when it becomes solid, and on this account it is used for taking casts from medals. When cast in a mold it expands into every line, and makes a most beautiful and faithful copy.

Bismuth, like a true metal, forms bases, and, like a non-metal also, it forms acids. But its metallic nature is much more distinct than that of antimony, since it is less likely to form acids. This is also shown by the fact that it has no compound with hydrogen.

**The Group.** — Bismuth and antimony are much alike, and, as we have just seen, both have the properties of non-metals. In fact, they agree closely with the nitrogen group. In bismuth the metallic properties are very clear, in antimony not so clear, in arsenic quite obscure, in phosphorus and nitrogen altogether wanting. From bismuth to nitrogen, the transition from metal to non-metal is gradual and perfect.

No better proof is needed that it is the chemist, and not nature, who divides the elements into these two great classes.

Such a division, however, is convenient. But chemists do not all agree as to just where the line shall be drawn. Some have put arsenic among the metals; others have put arsenic, and antimony too, among the non-metals.

In this book the line has been drawn between arsenic and antimony. Because, then, we have on one side, all the elements whose compounds with oxygen and hydrogen are *acids*, and on the other, all those whose compounds with oxygen and hydrogen are ever *bases*. The oxides of arsenic are always acid, while one of the oxides of antimony is basic, — feebly so to be sure, but truly basic.

COMPARE THEIR REACTIONS. — Some of the chemical actions of bismuth, antimony, and arsenic will be a useful study by experiment.

*Ex. 169.* — Arrange three tubes each with about 10 cc. of water. Into one put a few grains of arsenous oxide; shake it well, and if it does not dissolve, heat it. It dissolves to a clear solution.

Into a second tube put a few grains of bismuth nitrate; shake it well: you find the liquid changed to a milky whiteness, instead of a clear solution. To this then add strong hydrochloric acid, drop by drop, and it will soon become perfectly clear.

Into the third tube put a few drops of antimony chloride; this liquid also at once becomes white. But add hydrochloric acid as before, and it becomes clear.

The compounds of bismuth and antimony are decomposed by *water*, but of arsenic not. This effect of water is not common. The white solids, precipitated by water, are soluble in hydrochloric acid.

These three clear solutions may be used in the following experiments. We will first study the sulphides, as follows:

*Ex. 170.* — Fit up the apparatus for hydrogen sulphide, Ex. 107. Put an arsenic solution — enough to cover the end of the glass tube — into flask *a*, and add a few drops of H Cl. Put an antimony solution, made clear by H Cl, into flask *b*, and a bismuth solution, also made clear by H Cl, into flask *c*, and then pass the $H_2S$ gas through them all.

Observe that the sulphides of all three metals are produced, and notice their different colors.

These are the first sulphides which we have found to be insoluble in acid water. And of these, one, the orange-red antimony sulphide, is soluble if much H Cl is present.

*Ex. 171.* — Let the sulphides settle, and then pour off the liquid; add considerable water to each. Let the

sulphides settle, and decant the liquid again. We repeat this work in order to wash the sulphides clean: do it once more. Now pour *yellow* ammonium sulphide, $(NH_4)_2S$, upon each, and gently warm it. The question is, whether the sulphides are soluble in this liquid. The bismuth sulphide will remain unchanged, but the other two should disappear; they are soluble in yellow ammonium sulphide.

*Ex. 172.* — Next compare the hydroxides. For this purpose make three clear solutions, as in Ex. 169, only be more careful to not use more HCl than just enough to dissolve the white precipitates. Then add $NH_4HO$, little by little, to each. Observe the difference between the arsenic and the other two solutions.

QUERIES. — By what experiments could you decide whether a given substance is a compound of aluminum or of chromium?

By what experiments could you decide whether a given substance is a compound of aluminum or of iron?

If a given substance is a compound of aluminum, or else of a metal in the antimony group, by what experiment could you decide which it is?

SUGGESTIONS. — Take one or more specimens from the teacher or a friend who knows what they are, and decide, by experiments, whether each is, or is not, a compound of aluminum.

Take specimens, which may be compounds of any one of the metals in the antimony group, and, by experiments, decide which.

## TIN AND LEAD.

**Tin.** **Sn''.** — Tin seems to have been known since the earliest periods of history.

It was called *Jupiter* by the old alchemists, and *stan*[1] by the people of Phœnicia. The people, just named, discovered tin in Britain more than a thousand years before the Christain era.

Tin is found in the form of *tinstone.* Tinstone is a compound of tin and oxygen, $SnO_2$; it is the chief ore of the metal.

The ore of tin is found in Cornwall, England, which has been noted for its tin-mines for many hundred years. Bohemia, Saxony, and Malacca and Banca, in India, also yield the ore of this metal. In this country it seems to be very rare, but it has been found in New Hampshire and California.

Extraction of the Metal. — To obtain tin from this ore would be a simple thing if the ore were pure; for it is an oxide which can be reduced by carbon. But the ore is mixed with other matters, which make the process more difficult.

To get rid of the rocky and earthy parts, the ore is *stamped* to powder under wooden stampers shod with iron, and then washed with water. To get rid of sulphur and arsenic, the ore is *roasted* and washed, perhaps twice. And then, to take away the oxygen, the roasted ore is *melted* with charcoal and lime.

Properties of the Metal. — Tin is a silver-white metal with a fine luster, and somewhat harder than lead. It can be hammered into sheets so thin that a thousand

[1] The symbol for tin, Sn, comes from its old name *Stan*, or *Stannum.*

or more would be needed to make an inch in thickness. It does not tarnish readily in air, showing that its attraction for oxygen is slight, but when heated, tin and oxygen combine at once.

Because tin is so malleable it is much used in thin sheets, called *tin-foil.* But the tin-foil of commerce usually contains some lead, and sometimes a large proportion of that metal.

Because air does not readily tarnish it, tin is used for coating iron to keep it from rusting. The "tin" of which common tin-ware is made is sheet-iron with a thin coating of tin.

**Compounds of Tin.** — There are two compounds of oxygen and tin; one is called stannous oxide, $Sn'' O$, the other is the stannic oxide, $Sn'''' O_2$. There are also two chlorides, two sulphates, — in fact, there are two classes of tin compounds, the *ous* and the *ic* forms.

Starting with the metal itself we can make several of the tin compounds and study by experiment their chemical behavior. We will first study the chlorides.

*Ex. 173.* — Place 5 cc. strong hydrochloric acid in a test-tube and drop into it a piece of granulated tin. Then heat until the effervescence is brisk; after which keep the tube warm by holding it above the flame of the lamp until, when taken away from the heat, the bubbling nearly or quite stops. If, before this occurs, the tin is used up, another piece must be added. When the effervescence is brisk, a match-flame, brought to the mouth of the tube, will cause a slight explosion.

What gas is set free by the action?

Into what compound is the tin changed?

$$Sn'' + 2\,H\,Cl = Sn''Cl_2 + 2\,H.$$

The $Sn\,Cl_2$ is stannous chloride.

It is easy to change this *ous* chloride into the *ic* form just as ferrous was changed into ferric chloride, Ex. 157.

*Ex. 174.* — Pour about 2 cc. of the $SnCl_2$ solution into another tube, add four or five *drops* of strong nitric acid, and boil the mixture a minute.

Is there any evidence of a chemical change?

But we may go further and prove that stannous chloride is no longer present, in this way:

*Ex. 175.* — Prepare two clean tubes with 10 cc. of water in each, and add to one about a cubic centimeter of the solution of stannous chloride of Ex. 173, and to the other about as much of the solution just now made. Next add, drop by drop, a solution of mercuric chloride to the first.

What is the color of the precipitate by the first drop?

What change occurs when more and more are added?

These are the changes which this reagent always makes in stannous chloride. Now add, to the other solution, some drops of mercuric chloride in the same way. No such changes should take place. Then it is plain that stannous chloride is not present.

The fact is that the stannous chloride took the chlorine from the mercuric chloride, and so became stannic chloride. The white precipitate at first, was mercurous chloride, and the gray at last, was mercury itself. Thus:

1. $2\,HgCl_2$ (Mercuric chloride) + $SnCl_2$ (Stannous chloride) = $SnCl_4$ (Stannic chloride) + $Hg_2Cl_2$ (Mercurous chloride)
2. $Hg_2Cl_2$ + $SnCl_2$ = $SnCl_4$ + $Hg$

In this way mercuric chloride will always tell one whether a given solution of a compound of tin is stannous or stannic chloride.

We will next study the sulphides. The two sulphides may be made by hydrogen sulphide. Thus;

*Ex. 176.* — To 10 cc. of water add ½ cc. of the stannous chloride of Ex. 173. And again, to 10 cc. of water add ½ cc. of the stannic chloride of Ex. 174. Then pass hydrogen sulphide through both.

The first gives stannous sulphide, Sn S, brown.

The second gives stanic sulphide, $Sn S_2$, lemon yellow.

It is plain that tin must be added to the list of metals whose sulphides are insoluble in acid water.

Now try these sulphides with yellow $Am_2S$. Are they soluble in it?

Refer back to antimony, arsenic, and bismuth, to see which of these have sulphides, which behave in the same way, with $Am_2S$.

### LEAD. *Pb''*.

The lead ores of Spain and of England were worked by the ancient Romans, and still farther back, even before the sacred books of Exodus and Job were written, this metal was known and used.

The ores of lead are many; but for the most part they are not abundant. One of them, however, is found in immense beds and veins in the rocks. This ore is called *galena*. It is a compound of lead and sulphur, Pb S.[1]

Illinois, Iowa, and Missouri, to say nothing of several other of our States, have an abundance of galena stored away in their rocks.

**Extraction of Lead from its Ores.** —To obtain the metal, the ore is *roasted*, on the floor of a furnace, with plenty of air. The hot oxygen changes a part of the sulphide into sulphate and another part into oxide, while a third part remains as it was, — sulphide.

The furnace is then shut tight, and the fire driven to a

[1] The symbol for lead, Pb, is from the Latin name of the metal, *Plumbum*. Lead is bivalent, Pb''.

greater heat. The new compounds, just mentioned, then attack and decompose each other, and all three of them give up their metallic lead. The way in which these three compounds reduce each other may be seen by writing the reactions. The three substances made by roasting are $PbSO_4$, $PbO$, and $PbS$.

Then $PbSO_4$ with $PbS$ become $2Pb$ and $2SO_2$
and $2PbO$ " $PbS$ " $3Pb$ " $SO_2$

The sulphurous oxide is carried away by the draught, while the melted lead remains covered with the earthy impurities of the ore, — a melted slag, from under which it may be drawn off.

By Iron. — There is another way of getting the lead from galena, based on quite another principle. In fact it may be given to illustrate another *method* of metallurgy, p. 194. In this method one metal is used to liberate another. This is called the *precipitation process.*

Some lead may be easily obtained by "precipitation," and the experiment will illustrate the fact that one metal may displace another from its compounds. Thus:

*Ex. 177.* — Dissolve 8 g. or 10 g. of lead acetate, commonly called "sugar of lead," in about 500 cc., a pint of water, and if the solution is cloudy add a little acetic acid to clear it. Put this into a white glass bottle, and then hang in it a strip of clean sheet zinc (Fig. 66), and let it stand undisturbed. It will not be long before brilliant crystals of lead may be seen on the surface of the zinc, but it should be left until to-morrow, that we may witness the beautiful growth of crystals, which has long been called the *lead-tree.*



Fig. 66.

Lead acetate and zinc, become zinc acetate and lead. This shows the precipitation of a metal "in the wet way."

But in the case of lead ores, iron, instead of zinc, is used to liberate the lead, and the change is brought about by heat instead of in solution. The ore and scraps of iron are heated together in a blast-furnace, when the iron takes the sulphur away from the lead, thus:

$$PbS + Fe = FeS + Pb$$

This is the "precipitation" of a metal "by heat."

**Properties of Lead.** — Lead is so soft as to be easily cut with a knife. Its freshly-cut surface shows that the metal has a light-blue color and a fine luster, which may be seen, in the crystals, in Ex. 177. It is a heavy metal, being 11.4 times heavier than the same bulk of water.

Its attraction for moist oxygen is quite strong, even at common temperatures, so that its surface is never bright except when freshly cleaned. But why say "moist" oxygen? Because it is found that in perfectly dry air lead does not tarnish, and also that in water which contains no air it stays bright, which shows that both air and moisture are required.

**Lead Oxides.** — Lead combines with oxygen to make three oxides. When heated in air the metal is changed to lead oxide, PbO, — a yellow powder known as *litharge*, which is used in making flint-glass, p. 185.

There is also the lead dioxide, $PbO_2$, which is a brown powder, and then another oxide, $Pb_3O_4$, called *minium*, or *red lead*, which is used as a paint.

Many of the lead-salts are insoluble; in fact, the nitrate and the acetate are the only two, at all common, which will dissolve in water. From these two, the student can make other compounds of lead, and by so doing become acquainted with the chemical actions of this metal.

*Ex. 178.* — To make the carbonate. First make a dilute solution of lead acetate, and make it clear by acetic acid

if need be, but use just as little acid as will answer this purpose. Then add some ammonium carbonate. The white solid obtained is "lead carbonate."

But it is found that this lead carbonate is not a pure carbonate; it contains lead hydroxide also. It is often called the *basic carbonate.* This basic carbonate is the "white lead" which is so much used as a paint.

Salts which contain an hydroxide are called *basic salts.*

*Ex. 179.* — To a dilute solution of lead nitrate[1] add ammonium hydroxide. The white precipitate is not a pure hydroxide: it contains lead nitrate also. It is another basic salt, called the basic nitrate.

Of what other metals are the hydroxides white?

Is this lead precipitate soluble or insoluble in excess?

What other hydroxides are like it in this respect?

*Ex. 180.* — To a *dilute* solution of lead acetate add drops of hydrochloric acid. No precipitate should appear. But now use hydrogen sulphide. The black precipitate which falls is lead sulphide, Pb S.

What other metals have given sulphides by $H_2S$ in acid?

Which of these other sulphides does this lead sulphide resemble in color?

Will $Am_2S$ dissolve this lead sulphide?

*Ex. 181.* — To a *strong* solution of lead acetate add drops of hydrochloric acid. A white precipitate appears: it is lead chloride, $PbCl_2$. Now heat the mixture, and observe that the chloride disappears. Let it cool again, and see that the chloride returns in the form of needle-shaped crystals. Why did not drops of HCl make a precipitate of this chloride in Ex. 180, as well as in this one?

All this proves that lead chloride is somewhat soluble

[1] Make the lead nitrate by adding drops of $HNO_3$ to the basic carbonate of Ex. 178, until the white substance is dissolved.

in cold water (Ex. 180), but not freely (Ex. 181), while in hot water it dissolves largely (Ex. 181).

Have we found any other metal, so far, whose chloride is insoluble?

If a solution contains either Bi or Pb, can you tell which by Ex. 181?

If a solution contains Bi or Pb, can you tell which by Ex. 180?

*Ex. 182.* — To a moderately strong solution of lead nitrate add drops of potassium iodide. The bright yellow precipitate is lead iodide, which is very sparingly soluble in the cold.

But heat the mixture and the iodide disappears. If there is water enough it will become perfectly clear. Now let the tube and contents cool, and watch it.

Describe the iodide as it now appears.

Why does the iodide reappear?

Lead is the only metal whose compounds yield, in this way, such a brilliant and crystalline iodide. Hence this experiment is an excellent test for lead. The appearance of rich yellow lead chromate, by the use of potassium chromate (Ex. 163), is another excellent test for lead.

QUERY. — By what experiments could you decide whether a given compound is, or is not, a compound of tin or lead? Try it.

## THE COPPER GROUP.

### COPPER. *Cu″*.

COPPER is often found in nature in the metallic state. This is the case in the noted copper-mines of Cornwall and Devon, in England. Some of the finest native copper in the world is found in the region of Lake Superior, where it occurs in great abundance. One single mass, of Lake Superior native copper, weighed over 400 tons.

Native copper is crystalline. The separate crystals are usually little cubes, but in some cases the cubes are grown together in vast numbers making up quite large masses, and these masses often show most singular branch-like forms, sometimes rudely resembling the form of some growing plant. Fig. 67 is the picture of a specimen.



Fig. 67.

But copper, as native metal, is much less abundant in nature than are its ores.

Copper pyrites, made of copper,[1] iron and sulphur, $Cu\,Fe\,S_2$, is the ore which is most common. It is crystallized in cubes of perfect form, having the color and luster of brass. Besides this there are other sulphides, such as $Cu_2S$ and $Cu\,S$, also found in considerable quantity.

Malachite is a rich ore of copper, less common than pyrites. It is a green stone, which takes a fine polish, and is often used for ornamental purposes. In composition it

[1] The symbol for copper is Cu, from the Latin name of the metal, *Cuprum*.

is a basic carbonate, for it contains both the carbonate and the hydroxide of copper, $CuCO_3, Cu(HO)_2$.

Other ores of copper are widely distributed. Some are blue, some are red, some are purple, some are gray, but we need not stop to describe them.

**Extraction of Copper.** — The metal is extracted from its sulphide by roasting and reducing it in air. The ore is first roasted and then melted, and then roasted again; this changes a part of the sulphide into copper oxide, Cu O. This roasted ore is then mixed with sand and heated in a reverberatory furnace. The copper goes back into the form of sulphide while the iron of the ore takes oxygen, and, with the sand, becomes a liquid silicate. At the end of this repeated roasting the copper is still combined with sulphur, but it is rid of the iron.

The rest of the operation is like that of getting lead. The sulphide is again roasted; a part of it is changed to oxide, while the rest remains as sulphide, and then, by heating them strongly, these two compounds attack each other, and the copper of both is set free. Thus:

$$Cu_2S + 2CuO = 4Cu + SO_2$$

**Properties.** — This metal has a peculiar deep red color, not to be seen in any other. It is rather soft, easily bent, very ductile, and very strong. It is one of the very best conductors, and this makes copper, more than any other metal, useful in all the applications of electricity.

**Alloys.** — The alloys of copper are many and important. Brass is an alloy of copper and zinc; German silver, of copper, zinc, and nickel; while bronze, and bell-metal, and gun-metal, are made of different proportions of copper and tin.

**Copper Compounds.** — Copper, when long exposed to moist air, turns green; the green coating is a carbonate. The metal does not unite with oxygen at common temperatures of the air, but when heated it does, and if intensely

heated, it slowly burns, giving a fine green color to the flame which heats it. This may be easily shown by holding a small copper wire for some time in the edge of the Bunsen flame. There are two oxides of copper: one is the cuprous oxide, $Cu_2O$, which is red; the other is the cupric oxide, $CuO$, which is black. So there are also two chlorides, two iodides, two sulphides, and, in fact, two classes of copper compounds, the cupr*ous* and the cupr*ic*.

**Cupric Sulphate.** — The cupric sulphate is the most common salt of this metal. It is usually called copper sulphate. It comes in the form of blue crystals, $CuSO_4 + 5H_2O$. The blue color depends upon the water of crystallization, for when the water, $5H_2O$, is driven off by heat, as it may be very easily, the substance is white. This salt goes under the common name of "blue vitriol," just as the ferrous sulphate is called "green vitriol," and as the zinc sulphate is called "white vitriol."

**Study of Some Reactions of Copper.** — Starting with the copper sulphate the student can produce several of the compounds of copper, and in this way become acquainted with some of the chemical peculiarities of this metal.

*Ex. 183.* — Reduce a little copper sulphate to powder and dissolve it in water little by little. First see how little will be needed to give a perceptible blue color to, say 10 cc. of cold water in a tube. Then add more and more until it no longer disappears when shaken.

Many other copper compounds in solution have the same color as this one. Indeed, when this blue color is seen in a liquid it is a *sign* of the presence of some copper compound, — not a proof, but a sign. This solution may be used for the experiments which follow.

*Ex. 184.* — Add a cubic centimeter of the copper sulphate solution to 10 cc. of water. Into this put slowly, one, two,

three *drops* of ammonium hydroxide, and shake it well. The precipitate, which is deep blue when just enough of the reagent is used, is copper hydroxide, $Cu(HO)_2$. With too little reagent the precipitate is pale blue.

Is this hydroxide soluble in excess?

Compare this result with the effect of ammonia on zinc, nickel, and cobalt, hydroxides.

Compare it also with lead and bismuth hydroxides.

*Ex. 185.* — Add a cubic centimeter of the copper sulphate solution to 10 cc. of water, and a few drops of hydrochloric acid.

Note the evidence that copper chloride is soluble. Then add hydrogen sulphide: black Cu S is made.

What other metals have given sulphides in acid water?

Which of those others does this Cu S most resemble?

Find out whether it is soluble or insoluble in $Am_2S$.

The same brown-black sulphide is made when $Am_2S$ is added directly to the original solution of cupric sulphate:

$$CuSO_4 + (NH_4)_2S = CuS + (NH_4)_2SO_4$$

and ammonium sulphate $(NH_4)_2SO_4$ is also made, which remains dissolved.

*Ex. 186.* — Place a bright piece of iron wire, or a knife-blade, in a dilute solution of copper sulphate in water: it is soon coated with red metallic copper. Leave it in until the *blue* color is *all gone* from the liquid. Then try a fresh piece of bright iron. It will need some time, but the iron will take every particle of copper out, so that this fresh piece will not be coated at all.

$$CuSO_4 + Fe = Cu + FeSO_4$$

Now by this reaction we see that, besides the metallic

copper, Cu, some ferrous sulphate, $FeSO_4$, must be made at the same time.

Does the color of the liquid *suggest* the presence of an iron compound? See Ex. 155.

Can you *prove*, by experiment, that it does contain iron?

### MERCURY. *Hg″*.

Mercury[1] is sometimes found in the earth as native metal, but oftener in combination with sulphur. The sulphide, HgS, is its chief ore, and it is called *cinnabar*. The color of this ore is dark red, and in some specimens it is almost as rich and brilliant as vermilion. Indeed, these two things have the same composition; cinnabar is the native sulphide, vermilion is the artificial sulphide. The ore is found in many countries; Spain, Austria, China, and California are examples.

**Extraction of the Metal.** — Mercury is easily obtained from cinnabar. The ore only needs to be roasted, when this reaction will occur:

$$HgS + 2O = Hg + SO_2$$

The ore is decomposed by the hot air; its sulphur burns off as sulphurous oxide, while its mercury is left free. The metal is then in the form of vapor, which is led into cold vessels, where it becomes liquid.

PROPERTIES. — Mercury is a liquid, — the only liquid metal. It is tin-white, with a splendid luster. It is very heavy, — thirteen and a half times heavier than the same volume of water. When cooled down to —39.5° C. it freezes to a white lustrous solid; when heated to 350° C. it boils.

With some other metals it makes alloys, but the alloys of mercury are generally called *amalgams*. Zinc, copper,

[1] The symbol for mercury, Hg, is taken from the Latin *Hydrargyrum*.

silver, and gold are changed into amalgams at once by contact with mercury.

**Compounds of Mercury.** — Mercury does not tarnish in air unless heated, but if kept at almost boiling heat, it is changed into mercuric oxide, Hg O, known as the "red oxide," Ex. 5. This metal also forms two series of compounds, the *ic* and the *ous*. Thus we find mercuric chloride and mercurous chloride, mercuric and mercurous sulphates.

Mercuric chloride, $HgCl_2$, is the virulent poison, known under the name of *corrosive sublimate.*

Mercurous chloride, $Hg_2Cl_2$, is the medicine known under the name of *calomel.*

Both of these are white solids, and when seen in powder are alike in appearance, but the poison is soluble in water, while the medicine is insoluble.

MERCUROUS COMPOUNDS. — The metal mercury is soluble in nitric acid, which, if dilute, changes it to mercurous nitrate.

*Ex. 187.* — Put a small globule of mercury into a tube, and add a cubic centimeter of dilute nitric acid. Then heat it. The mercury will slowly waste away, and, if need be, let more acid be added, but not too much, so that the metal shall be completely dissolved.

This solution of mercurous nitrate may be used as follows: —

*Ex. 188.* — Add a few drops of it to 5 cc. of water, and then add drops of hydrochloric acid. The white solid obtained is mercurous chloride, or calomel, $Hg_2Cl_2$.

What other metal yields a precipitate with H Cl?

Is this chloride, like that one, soluble by boiling?

Now add to the white chloride a little ammonium hydroxide and note the marked change in color. This is

the only white chloride which will turn black when treated with ammonia.

*Ex. 189.* — Add a few drops of the nitrate to 5 cc. water, and then add ammonium hydroxide. The same black precipitate is made at once.

Mercurous compounds are the only ones which yield black precipitates with ammonia.

MERCURIC COMPOUNDS. — A solution of corrosive sublimate, mercuric chloride, may be the starting-point in our study of the chemical actions of mercuric compounds.

*Ex. 190.* — Add a cubic centimeter of mercuric chloride to 5 cc. of water. Add also drops of hydrochloric acid. No precipitate will appear.

Can you explain the fact that H Cl can make no precipitate in *any* mercur*ic* compound?

Finally add hydrogen sulphide, and notice any changes of color which may occur. At first the precipitate may be white, but with more and more $H_2S$ it will become yellow, brown, and at last black. The mercuric sulphide, Hg S, when well formed is black.

What other metals have given black sulphides by $H_2S$?

*Ex. 191.* — Add a cubic centimeter of $HgCl_2$ to 5 cc. of water, and then use $NH_4HO$, little by little, to excess.

What is the difference between the actions of $NH_4HO$ on the mercurous and the mercuric compounds?

*Ex. 192.* — To another dilute solution of the $HgCl_2$ add one drop of hydrochloric acid, and then put into it a piece of bright copper wire.

Notice the coating which soon gathers on the wire.

What is it? Its appearance, especially when rubbed, will be likely to inform you. So will the reaction if it be written, thus:

$$HgCl_2 + Cu = CuCl_2 + Hg.$$

### SILVER. *Ag'*.

Silver is sometimes found in the native state. Such specimens are often very beautiful, looking like metallic twigs and branches. But this metal is more abundant in combination, sometimes with chlorine, but oftener with sulphur. The sulphide of silver is sometimes found alone, but, in most cases, it is in company with the sulphides of other metals, such as lead and copper. These mixed sulphides of silver and other metals are the chief ores of silver.

Silver is almost always present in galena, and much silver is obtained from that ore.

**Extraction of the Metal.** — There are different ways of taking silver from its ores. We will study two of these.

First suppose the silver is to be taken from its sulphide, which contains the sulphides of other metals, as it usually does. Can the ore be treated the same as the ores of zinc or lead, — that is, changed to an oxide by roasting in air? No, because silver, unlike those metals, has very little attraction for oxygen. But it has a very strong attraction for chlorine, and this fact is made use of to get the metal away from the sulphur in its ore.

For this purpose, the ore is well stamped, to crush it to powder, and then mixed with a little — about one tenth its weight, common salt — sodium *chloride*. The mixture is then heated, and, to make sure that the work is well done, the ore is mixed with more salt and heated again. In this operation, the chlorine of the salt seizes the silver of the ore, and holds it in the form of silver chloride.

The next thing is to take the chlorine away, and iron is used for this purpose because chlorine has a stronger

attraction for iron than for silver. The chloride of silver is put into strong oaken casks, with a little water and fragments of iron. This mixture is violently shaken by turning the casks rapidly, and, under this treatment, the iron robs the silver of its chlorine. The silver is thus left free, but it is in very fine particles scattered through the mixture in the casks. How shall it be collected?

Just at this point the use of mercury comes in. Mercury is poured into the casks, which are then set rotating again faster than before. This brings the silver and mercury together, and they at once combine to form an amalgam, p. 251, and the next thing is to get the silver out of this amalgam.

To do this it is only necessary to heat the amalgam, when the mercury will pass off as vapor while the silver will be left behind.

But the metal is still mixed with such other metals, — copper, iron, and antimony, as were in the ore. These must be removed, and so the metal is mixed with lead, and again melted while a current of air is driven over its surface. The oxygen seizes the lead and other metals, but not the silver, which is thus left pure.

SILVER FROM GALENA. — But much of the silver now in use comes from the lead ore, galena. When the lead is obtained from this ore by the method described, p. 242, the silver comes with it, and then the problem is, how to get the small quantity of silver out of the large mass of lead.

The lead is melted, and then allowed to slowly become cold. When it cools, crystals begin to form, and these crystals are pure lead. These crystals of lead are taken out with a strainer. They are melted over again, and again cooled; another crop of pure lead crystals is then taken out, and this heating, and cooling, and dipping out

pure lead is continued, the silver every time remaining in the molten mass, until what is left behind is quite rich in the precious metal.

And then oxygen is called upon to take away the lead that remains. The rich alloy is melted on a bed of bone-ash, and hot air is driven over it. The hot oxygen seizes the lead, but does not combine with the silver. The lead oxide soaks into the porous bone-ash and leaves the pure silver behind. This method of separating silver from lead is called *cupellation.*

**Properties of Silver.** — Silver is the whitest of metals. It is ten and a half times as heavy as water. It is much harder than gold, and very malleable and ductile.

Oxygen does not attack it even when hot; but its surface will quickly blacken in the presence of hydrogen sulphide because silver and sulphur strongly attract each other.

Articles of silver-ware and silver coin are well known. But we must remember that none of these are made of pure silver. The pure metal is quite too soft to be able to stand the wear which these things receive. To harden the silver, small portions of other metals are mixed with it.

For silver coin, some copper is added to harden the precious metal. In this country the standard coin metal is made of silver 90 parts and copper 10 parts. In England the proportion of copper is less (7.5), in Germany it is more (12.5).

**Compounds of Silver.** — The most important compound of silver[1] is the nitrate, $AgNO_3$, and this is made by the action of nitric acid on the metal. This action begins instantly on putting the two together, and goes on vigorously until one or the other is exhausted.

$$1.\quad Ag + HNO_3 = AgNO_3 + H$$
$$2.\quad H + HNO_3 = H_2O + NO_2$$

[1] The symbol for silver is Ag, from the Latin name *Argentum.*

The silver takes the place of hydrogen in the acid, making the silver nitrate, $AgNO_3$, and setting the hydrogen free, but this hydrogen instantly attacks another portion of nitric acid, and forms water and nitrogen dioxide. The liquid left is a solution of silver nitrate. By evaporating this liquid, transparent crystals of the silver nitrate may be obtained.

*Ex. 193.* — To 10 cc. of water add 4 or 5 drops of silver nitrate, and then add drops of hydrochloric acid. A white precipitate forms at once: what is it?

Boil this precipitate, as lead chloride was boiled in Ex. 181. Is it soluble in hot water, like lead chloride?

Add ammonium hydroxide to it, as was done with the mercur*ous* chloride, Ex. 189.

Does it blacken by ammonia, as the mercurous chloride did? What is the effect of the ammonia on it?

By these differences, we can easily distinguish these three chlorides, one from another.

*Ex. 194.* — To 10 cc. water add drops of silver nitrate, and then hydrochloric acid, or a solution of common salt, which will have the same effect. Place the tube with this white silver chloride in sunlight, and notice its change of color.

Other silver salts are also very sensitive to the action of light. On this account the compounds of silver are much used in the art of *photography.*

**Separation of Silver and Copper.** — We have seen that silver coin is an alloy of silver and copper. The two metals have been melted together: can we get them apart again, or prove that they are in the coin? The student is now able to do this, and the work is an interesting and instructive exercise.

*Ex. 195.* — Put a ten-cent piece in an evaporating-dish and cover it with a mixture of strong nitric acid and water,

half and half. Then heat gently, until chemical action begins. Describe the action, thus:

What gas is set free? What color does the liquid become?

What compound of silver is made? What compound of copper?

When the chemical action is over, or nearly so, pour about 1 cc. of the blue liquid into about 5 cc. of water, and then add drops of hydrochloric acid, or, better, of solution of common salt, Na Cl, until a drop no longer makes a precipitate. Let the precipitate settle, and then pour the liquid off into another tube.

Wash the precipitate by pouring water upon it, shaking, and then, when it is settled again, pouring the water off into the waste. Do this a second time.

Now you have the copper of the coin in the blue liquid, and the silver of the coin in the white precipitate.

Prove that the blue liquid contains *copper*, by the use of iron, Ex. 186.

Prove that the white precipitate is *silver* chloride by treating a part of it with hot water, another part with ammonia, and exposing another part to the sunlight.

SUGGESTION. — Take specimens from the teacher or a friend who knows what they are, and, by experiment, decide whether each one is, or is not, a compound of lead, of silver, or of mercury. If it prove to be a compound of mercury, then decide whether it is a mercuric or mercurous compound.

## GOLD AND PLATINUM.

### GOLD. $Au'''$.

Gold is almost always found in the native state. Unlike the metals described before, this precious metal is very rarely found in combination with other elements: it is found as native metal. Yet this native gold is not pure; it is mixed with silver, and sometimes, too, with copper, and often with other baser metals. Native gold is found in fine grains among the sands of certain rivers, and also in solid quartz rocks. Now and then a nugget of considerable size occurs.

**Extraction of Gold.** — Gold is obtained from sands and other loose materials by the process called "washing." The material is put into a shallow pan and well stirred up with water. Gold is so heavy that the grains will quickly settle to the bottom, and then the earthy matter may be poured off from above it. Sometimes the gold-bearing deposit is washed by rocking it in a cradle through which a stream of water is slowly running. The lighter earth or sand is then washed away, while the heavy *gold dust* lags behind, and is caught in grooves in the bottom of the cradle. The precious metal is then dissolved in mercury, and afterward separated by heat.

Gold is obtained from quartz rock by "amalgamation." The quartz rocks are crushed to the finest powder and then mixed with mercury. The mercury dissolves the gold and leaves the quartz. The amalgam of gold is then distilled, and the mercury goes away as vapor while the gold is left behind.

Properties. — Gold is remarkable for its fine yellow color and beautiful luster. It is among the heaviest of

metals, about nineteen times (19.33) heavier than water. It is the most malleable of metals; it is said that leaves have been beaten so thin that 280,000 would be needed to make an inch in thickness! There is a curious fact about the color of gold-leaf; it is this: looked at in the usual way it is yellow, but let a leaf be spread upon glass, so that it may be held up between the eye and a window, or the sky, and it will be green.

At a temperature of about 2000° F. (1102° C.) this precious metal melts into a greenish-colored liquid. The highest heat of a furnace can scarcely change it into vapor, but the furious flame of the oxyhydrogen blowpipe can; its vapor has a purple hue.

Few chemicals, even of the most corrosive character, can harm this noble metal. Oxygen cannot rust it; sulphur cannot tarnish it; nor can the strongest acids corrode it. For one element, however, it has a strong attraction; this is chlorine. On this account it dissolves readily in *aqua regia;* in fact this liquid received its name, which means royal water, because it was found to dissolve gold, the "king of metals." It changes the gold to gold chloride, $AuCl_3$.[1]

Gold is used for ornament and as money. But when pure it is almost as soft as lead, and hence unfit for either use. To make it harder a little copper or silver is added. The alloy used for coin in this country must be made of *nine* parts of gold to *one* part of copper.

### PLATINUM. *Pt.*

Platinum is a still rarer metal than gold. In small quantities gold is very widely distributed in the earth, but platinum is found in few places. It is found in largest quantity on the slopes of the Ural mountains, and

[1] The symbol for gold is Au, from the Latin *Aurum*.

in Brazil and Peru. It is always found in the native state, but mixed with other metals, several of which, like osmium and iridium, are still rarer metals than itself.

Pure platinum is silver-white. Oxygen cannot attack it at any temperature, and only the intense heat of the blowpipe can melt it. On these accounts platinum vessels are much used in the laboratory. Aqua regia and the caustic alkalies will dissolve platinum, and many other metals will form alloys with it, which melt more easily than the metal itself and are attacked by acids.

**The Platinum Group.** — There are five very rare metals which are much like platinum in many of their properties. They are palladium, iridium, osmium, ruthenium, and rhodium. Palladium is remarkable for its action on hydrogen. When heated to redness it absorbs this gas, and forms what seems to be a true *alloy*. And this fact would indicate that hydrogen is itself a metal. Accordingly it is sometimes called *hydrogenium*.

Osmium is remarkable for its great weight; it is the heaviest substance known, having a specific gravity of 22.477. An alloy of osmium and iridium is found sometimes in the sands with gold; it is harder than steel, and is used for the tips of gold pens.

## CLASSIFICATION OF THE METALS.

**How Classes are made.** — We have seen that some elements combine with oxygen and hydrogen to produce *acids* while others combine with oxygen and hydrogen to produce *bases*. If we put *all* those which produce acids together, and *all* those which produce bases together, we have all the elements in two large groups, — the non-metals and the metals.

Now this illustrates the way in which *classes* are made. If the chemist has a large number of substances, he may find, by studying them separately, that several are alike in some important character; these he can put together to form a class or group. Others in the list will not agree with these, but while unlike them they will be like one another, and these he can put together to form a second class or group. And in this way he may at length have the entire list divided into groups, — each group holding only those individuals which are alike in the leading character, however much they may differ in other particulars.

The Classes of Non-Metals. — We have seen that the non-metals are placed in four groups, known as the chlorine group, the sulphur group, the nitrogen group, and the carbon group. The leading character, or, we may say, the foundation of this grouping, is *valence*. All univalent non-metals are put into one group, all bivalent non-metals into another, all trivalent non-metals into a third, and the quadrivalent non-metals into a fourth.

The Metals not Classed in the Same Way. — Now the classification of the metals is also sometimes founded on valence. Then, all univalent metals go together into one group, all bivalent into another, and so on. But in

this book, the metals have not been grouped in this way, and the reason is, simply, that such groups of the metals would not best suit our particular purpose in the study of these elements. What we have wanted to do is to become acquainted with the general properties and behavior of the metals, and if they are grouped by their valence, metals quite different in properties and actions will sometimes be thrown together. As, for instance, silver is univalent, so are potassium and sodium, and all these will be in the univalent group if our classes are made by valence. But silver has few properties like the others. If we are studying the properties of the metals, it is better to put those together which are most nearly alike, and then silver will go with copper and mercury, instead of with potassium and sodium.

MORE THAN ONE WAY TO GROUP THEM. — In fact, the classes of metals in this book are founded on *likeness of properties*, and not on valence. Thus we have the potassium group, because there are several metals which are all very much like potassium, — soft, light, alkaline; and the calcium group, because there are several metals like calcium; and the iron group, because the properties of iron are much the same as those of the other members. This is the most helpful way of grouping the metals when our object is to study their properties.

But there must be some reason *why* some elements are so much alike and others so different. There must be some reason *why* some are univalent, others bivalent, and others with still different degrees of valence. If we could only get at this *cause* of their differences, the elements would fall into natural groups. We could then know the place of every one of them in the true system of nature. But the chemist does not yet know this *cause*. He knows the *facts* only. On this account the true system of the

elements in nature is still beyond his reach. And until he reaches it, the elements will be classified in different ways, as different leading facts are chosen on which the classes may be founded.

FOUR WAYS TO CLASSIFY THE METALS. — There are, in fact, four important ways of classifying the metals. One way is founded on *valence;* a second way is founded on *likeness of properties;* a third way is founded on *atomic weights;* a fourth way is founded on the *solubility of compounds.*

We need not stop longer with the first two of these; the third and fourth are yet to be described. The third is important because it includes all elements, metals and non-metals alike, in one system, and comes nearer than any other to being the "natural" system. The fourth is important because it is used in chemical analysis.

## THE NATURAL SYSTEM.

**Classification by Atomic Weights.** — We have seen that the elements chlorine, bromine, iodine, and fluorine are very much alike in properties. They also combine with the same elements and in the same proportions. But while they have many of the same properties they have them in different degrees. Fluorine is the most active, chlorine next, then bromine, and finally iodine is least active of all. Their *properties* vary in this order. And we have also noticed the fact that their *atomic weights* vary in the same order. The same fact has been noticed in our study of other groups.

NEWLAND'S DISCOVERY. — Such facts led Mr. Newland in 1863 to arrange many of the elements in the order of their atomic weights, beginning with the one whose atomic weight is least, and then, on looking over the list, he discovered that the elements with like properties were scattered at about equal distances apart.

Let us leave out hydrogen and begin with lithium, whose atomic weight is next larger than that of hydrogen, and put the atomic weights under the symbols of the elements in the order of value, and we have

| Li | Be | B | C | N | O | F | Na |
|---|---|---|---|---|---|---|---|
| 7 | 9 | 11 | 12 | 14 | 16 | 19 | 23 |

But sodium, Na, is like lithium, Li. Both are metals of the alkalies, and univalent. Now begin with sodium and go on:

| Na | Mg | Al | Si | P | S | Cl | K |
|---|---|---|---|---|---|---|---|
| 23 | 24 | 27 | 28 | 31 | 32 | 35.5 | 39 |

But potassium, K, is like sodium and lithium, alkaline and univalent. These properties are found in every *seventh* metal in the list so far, and not in any between.

Let us write the list, of the first twenty-one, in lines of seven. Thus:

| | | | | | | |
|---|---|---|---|---|---|---|
| Li = 7 | Be = 9 | B = 11 | C = 12 | N = 14 | O = 16 | F = 19 |
| Na = 23 | Mg = 24 | Al = 27 | Si = 28 | P = 31 | S = 32 | Cl = 35.5 |
| K = 39 | Ca = 40 | Sc = 44 | Ti = 48 | V = 51 | Cr = 52 | Mn = 55 |

And we find that the elements which are most alike fall together, under one another, as lithium, sodium, and potassium in the first column, and fluorine and chlorine in the last.

It is a curious fact that if we add 16 to the atomic weight of lithium, we get that of sodium, with the same properties. Add 16 again, and we get the atomic weight of potassium, another element with the same properties. Add about 16 to the atomic weight of any one and we get the atomic weight of another which has similar properties. This is true until we reach phosphorus, when we must begin to add 20 instead of 16.

MENDELEJEFF'S TABLE. — Now the Russian chemist, Mendelejeff, went so far as to make a table of all the known elements in the order of their atomic weights, and he found that all those elements, which are most alike, then fell into the same column. The horizontal lines of elements in this table are called "*series*" and the columns are called "*groups.*" In all there are 12 series and 8 groups. Three of these series are given on p. 265, and in them are the first three members of seven groups. The fact is that —

If all the elements are placed in the order of their atomic weights those which are most alike will be found at regular distances apart. As the atomic weights increase, the same properties appear, again and again, at regular intervals. This is known as the *periodic law.*

**The Spiral of Elements.** — The table of elements has been made in many shapes. The spiral form,[1] Fig. 68, shows the main point very clearly, which is, that the elements fall naturally into groups as their atomic weights vary by regular additions.

We begin with lithium at the center, and wind our way around the spiral, to the left, until we come out at the circumference. Along the way we pass by elements with larger and larger atomic weights in regular order. The spiral is divided into eight sections by lines from the center, and the elements which are alike in properties are found together in these sections. In fact, this continuous spiral of atomic weights throws all the elements into eight natural groups.

Take, for example, our chlorine group, p. 147, whose members have been found to be so much alike. We find them

[1] First given by Baumhauer in 1870, then modified by Huth in 1884, and adopted by Carnelly in 1885. — *Chemical News,* Vol. 53, p. 184.

all in the seventh group of the spiral, brought there by the order of their atomic weights. We find all the members of the oxygen group in the sixth, all of the nitrogen group in the fifth, and all of the calcium group in the second.

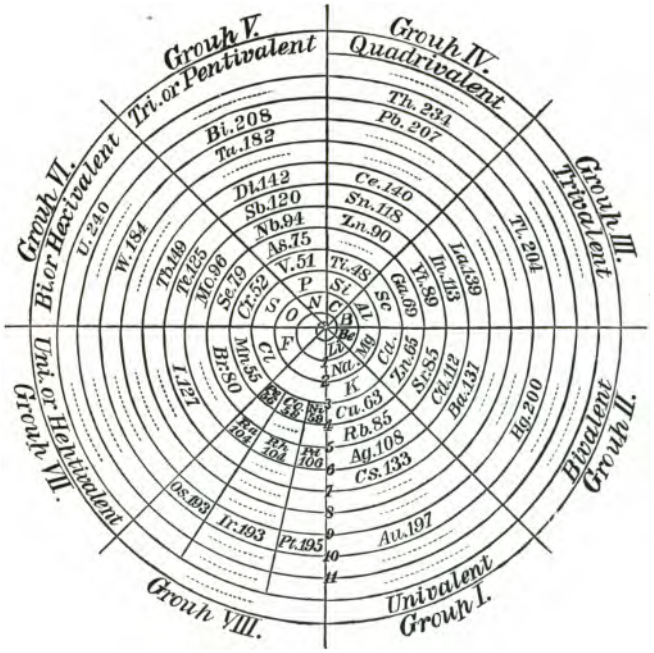



Fig. 68.

But in each of these "natural groups" in the spiral we find some elements which have not before been put together. In the first, for example, we find all the metals of the potassium class, and with them, copper and silver. Now many of the leading properties of copper and silver are not like those of sodium and potassium. They do not

agree with these in being soft, light, and combustible on water. Yet in some things they do agree quite closely. For example:

| | | | |
|---|---|---|---|
| Potassium | forms the | oxide | $K_2O$ |
| Copper | " | " | $Cu_2O$ |
| Silver | " | " | $Ag_2O$ |

In these compounds the atoms of potassium, copper, and silver hold the same relation to oxygen. And the order of their atomic weights brings them together in the same group in the spiral.

THE VACANT PLACES. — There are many vacant places in the spiral: what is the meaning of these gaps? They lead us to think that the elements are not yet all discovered.

When the table was first made there was a vacant place where gallium, Ga, now stands in the third group. Mendelejeff said that there ought to be an element whose "atomic weight is about 69, its specific gravity about 6, and its atomic volume about 11.5," to fill this gap. *Afterward gallium was discovered.* Its atomic weight proved to be 69.8, its specific gravity 5.9, and its atomic volume 11.8. From its vacant *place in the table,* Mendelejeff was able to give the properties of this metal before it had been discovered. He did the same thing for the element scandium, Sc, also in the third group. Possibly other elements are yet to be discovered by which other gaps will be filled.

### THE ANALYTICAL SYSTEM.

**Classification Founded on Solubility.** — We have found that hydrochloric acid will give a precipitate if mixed with a solution of a compound of silver, Ex. 193, but none if mixed with a solution of a compound of copper, Ex. 185. This is because silver chloride is insoluble in the liquid used, while copper chloride is soluble.

Now all metals whose chlorides are insoluble, like that of silver, can be put with silver to form one group. Such differences in solubility throw the metals together into classes called the *analytical groups.*

But we have found, by experiment, what many of the insoluble compounds of the metals are, and we can see, from our experiments, which are alike and which not alike. Indeed, if we will *just gather up the facts* which are scattered among our experiments, we shall find the metals falling into the analytical groups. If the student has made the experiments he will find it an interesting and profitable exercise to deduce this classification from them.

To do this make a table, with six columns, like the blank form below, and then fill the blank spaces opposite the numbers in each column, by consulting the record of the results obtained by the experiments with the metals, thus:

| CHLORIDES insoluble in acid water. | SULPHIDES insoluble in acid water. | HYDROXIDES insoluble in $NH_4HO$, and $NH_4Cl$. | SULPHIDES insoluble in $NH_4HO$. | CARBONATES insoluble in $NH_4HO$, and $NH_4Cl$. | ALL SALTS soluble. |
|---|---|---|---|---|---|
| 1. . . . . | 1. . . . . | 1. . . . . | 1. . . . . | 1. . . . . | 1. . . . . |
| 2. . . . . | 2. . . . . | 2. . . . . | 2. . . . . | 2. . . . . | 2. . . . . |
| 3. . . . . | 3. . . . . | 3. . . . . | 3. . . . . | 3. . . . . | 3. . . . . |
| | 4. . . . . | | 4. . . . . | | 4. . . . . |
| | 5. . . . . | | | | |
| | 6. . . . . | | | | |
| | 7. . . . . | | | | |
| Group I. precipitated by HCl. | Group II. precipitated by $H_2S$. | Group III. precipitated by $NH_4HO$. | Group IV. precipitated by $(NH_4)_2S$. | Group V. precipitated by $(NH_4)_2CO_3$. | Group VI. not precipitated. |

Place in the first column a list of all the metals which will give a precipitate with hydrochloric acid: you find their names by looking back over the experiments under

each metal. This list will contain the metals whose *chlorides* are insoluble in acid water. These metals make up Group I.

Place in the second column a list of all *the other* metals whose *sulphides* are insoluble in acid water. These will make Group II.

Place in the third, a list of all metals not found in the first and second columns whose *hydroxides* are not soluble in $NH_4HO$, or $NH_4Cl$. In this way we make out Group III.

Place in the fourth, a list of all the metals not found in the first three columns, whose *sulphides* are insoluble in $NH_4HO$. These are the metals of Group IV.

Place in the fifth, all the remaining metals whose *carbonates* are insoluble in $NH_4HO$, or $NH_4Cl$. These are the metals in Group V.

Place in the sixth, all the remaining metals whose chlorides, sulphides, hydroxides, and carbonates are all *soluble*. These are the metals in Group VI.

We now find that our study of the metals, by experiment, has not only given us the facts about them and their compounds, but that it has also given us a way to detect their presence in substances whose composition is unknown to us. In fact, we are now prepared to analyze a salt, thus: —

**To find out what Metal a Salt contains.** — If the salt is in the solid form we begin by making a solution of it. Then, —

1. — Taking a small quantity of the solution in a tube, we add hydrochloric acid drop by drop. If a precipitate is made in this way, it shows that the metal of the salt is one or another of the three in Group I. And then, to learn which one of these it is, we may try, with the solution, the experiments which are given under the

head of each of these particular metals, in our previous study of them.

But if the hydrochloric acid yields no precipitate, we may say that the salt is not a compound of any one of these three metals, and may seek further, to find whether it is one of the seven metals of Group II.

2. — The question is, whether its sulphide is insoluble in acid water. And to answer it, we take a little of the solution, make it acid by adding a drop or two of hydrochloric acid, and then add hydrogen sulphide, $H_2S$. If a precipitate is made in this way, our salt is a compound of one of the seven metals of Group II., but if none, it is not.

Now the sulphides of three of these seven metals can be dissolved by ammonium sulphide $(NH_4)_2S$, while those of the other four cannot. So we may take a little of our precipitate, and warm it with ammonium sulphide, and, if it proves to be soluble, we may say that the metal in our salt is one of these three, but if insoluble, we say that it is one of the other four.

In either case we may find out which particular one of the metals is present, by making the experiments given in this book in the description of each of them.

But if hydrogen sulphide has given no precipitate, we say that our salt is not a compound of any one of these seven metals, and go on to learn whether it is of one of the three in Group III.

3. — The question now is, whether its hydroxide is insoluble in ammonia and ammonium chloride. And to answer this, we take a little of our first solution, put into it, first a considerable quantity of ammonium chloride, $NH_4Cl$, and then add ammonia, $NH_4HO$, until the liquid smells strongly of this substance. It must be well shaken and the air blown out of the tube, then the odor of ammo-

nia will show when enough has been added. If a precipitate is made in this way, we may say that our salt is a compound of one of the three metals in Group III. And then, to know which one, we note the color of our precipitate and also make the experiments already described in the study of these three metals.

But if we get no precipitate, we say that our salt cannot be a compound of any one of these metals.

4. — The next question is, whether the sulphide of the metal we are after is insoluble in ammonia water. And to answer this, we take a little of our first solution, and put a little ammonia with it, then add ammonium sulphide drop by drop. A precipitate now shows that our salt is a compound of one of the four metals in Group IV. We then decide which one of the four, by noting the color of the sulphide, and also making the experiments given in our description of these metals.

No precipitate here, with ammonium sulphide, proves that no one of these four metals is present in our salt.

5. — We next ask, whether the carbonate of our metal is insoluble in ammonia and ammonium chloride, for if it is, then our metal is one of the four in Group V.

To a little of the solution of our salt we add considerable ammonium chloride, $NH_4Cl$, and then add ammonium carbonate $(NH_4)_2CO_3$. If we get a precipitate, we next decide which one of the four metals of Group V. is present in our salt, by the experiments already given in the study of these metals.

But if no precipitate is made, these four metals are thrown out of the question, and our salt must then be a compound of one of the metals in Group VI.

6. — Which one of these four is it? This is the final question. To answer it, we may make the experiments which are given in the descriptions of these metals, and see with which of the four our results agree.

By following the order of the groups in this regular way, one can bring out the metal of almost any simple salt, with certainty and very quickly.

The student may do this with a few simple salts whose names he does not know. The teacher will select them and keep a record of their names, and at the end of his work the student should report his work and the result of it. He should write down, in a short way, every experiment he makes, and what the result of it is, and what it proves; and he should do this for each experiment, *at the time he makes it*, and not wait until others are made.

On page 275 is a copy of one report of an "analysis," which shows a form of the notes to be kept in each case. Notice that, below the heading, which is the name of the student, date of the work, and number of the substance, the sheet is divided into three columns. In one, is put a brief description of all the experiments made. In the second, the result of each is written, and in the third, the fact which it proves.

**To find out what Acid the Salt contains.** — But when we have found the metal in a salt, we do not yet know the name of the salt. It may be a nitrate, a chloride, or some other compound. The next step is to identify the acid part of it. This is done by making the tests which were described in our study of the non-metals, or the experiments described in the *Exercises*. Thus a test for sulphuric acid is given on p. 167, and the use of it, to show whether a substance is a sulphate, is illustrated in Ex. 122. See also the Exercises, p. 173. Hydrochloric acid, or its salts, — the chlorides, — may be identified by the test on p. 147. See also the exercises on p. 151. Nitric acid, and the nitrates also, may be detected by the "Copperas Test," p. 100. If a salt is a carbonate

it will effervesce when treated with dilute acid, and the gas which is set free will whiten lime-water.

In this work for the acid, as in that for the metal, brief notes of all the experiments should be made, and reported.

**To Name the Salt.** — Having found what metal a salt contains, and also what the acid part of it is, you can declare the name of the salt which was given you for analysis (p. 136). For example, having found strontium and nitric acid, the substance itself must be strontium nitrate.

In this way, a "simple" salt, that is, a salt which contains only one acid and one metal, may be analyzed. The work is a most valuable exercise for the student at this stage of his progress, and, with a good reference book at hand, it should be carried as far as time will permit.

**Hint as to further Work.** — Even "complex" substances, that is, substances which contain more than one metal or acid, may be analyzed by very natural additions to the foregoing work. If, for example, we have a mixture of, say, silver and copper nitrates, we can detect both the silver and the copper. For we can separate them by using H Cl, as was done in Ex. 195.

If our mixture should contain one metal of every group, still every one of them could be "separated," and then "identified," by just such work. In such a case we could use the *group reagents* (see bottom of table, p. 269) one after another, in regular order. H Cl would take out *all* of the metal of Group I., and leave all the others in solution. Every group reagent in its order would precipitate *all* of the metal of its group, and leave all those of groups that come after in solution. In this way we should get each metal alone, and could then identify it. Of course, if a group reagent gives no precipitate, all metals of its group must be absent.

**REPORT OF AN ANALYSIS.**

*Name*_____. *Date*_____.

*Substance No.*_____.

| EXPERIMENTS. | RESULTS. | INFERENCES. |
|---|---|---|
| I. Add drops of H Cl. | No precipitate made | Absence of Group I. Ag, Hg (ous), and Pb. |
| II. To I. add $H_2S$. | No precipitate made | Absence of Group II. Hg (ic), Bi, Cu, Cd, As, Sb, Sn. |
| III. To the original solution add $NH_4Cl$ and $NH_4HO$. | No precipitate made | Absence of Group III. Fe, Al, Cr, |
| IV. To III. add $(NH_4)_2S$. | No precipitate made | Absence of Group IV. Mn, Zn, Ni, Co. |
| V. To original solution add $NH_4Cl$ and $(NH_4)_2CO_3$. | A white precipitate | Presence of Group V. Ba, Sr, or Ca. |
| VI. To the original solution add solution of $CaSO_4$. | No precipitate made in the cold | Absence of Ba. |
| VII. Heat VI. to boiling. | A white precipitate | Presence of Sr. |
| VIII. Flame-test. | Brilliant crimson | Confirms presence of Sr. |

Hence substance No._____ is a compound of strontium.

# APPENDIX.

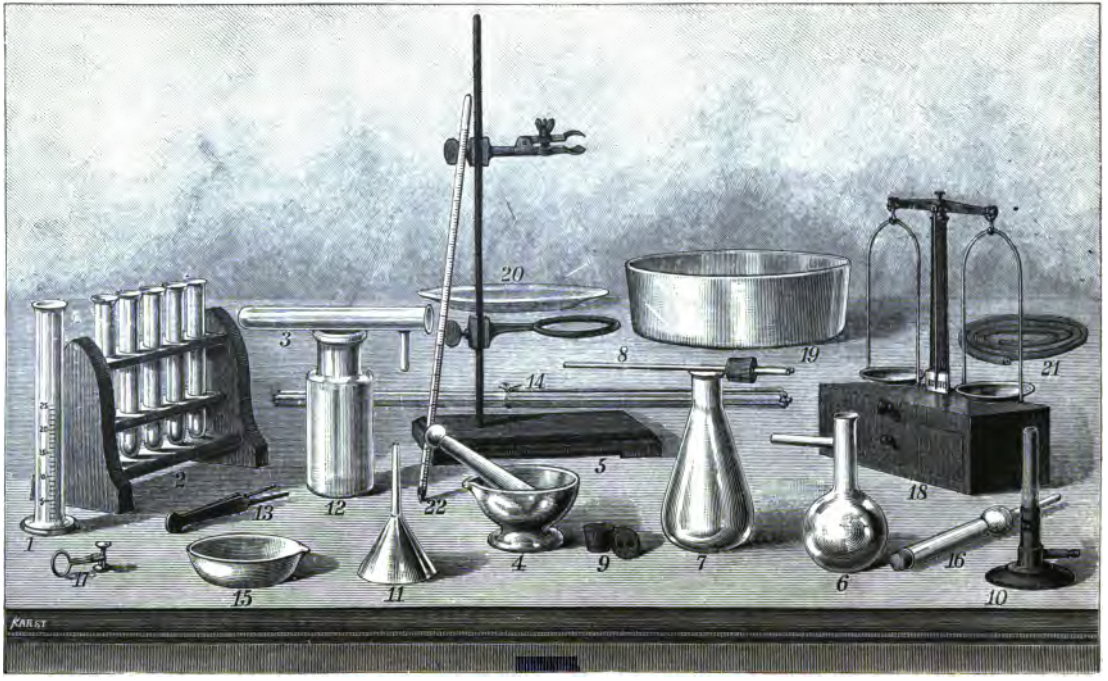



Fig. 69.

## THE APPARATUS.

THE following list contains the pieces in a single set of apparatus for the course of experiments described in this book. The full set is shown in the cut, Fig. 69. It is designed for the use of beginners, unused to manipulation, and of teachers who are oftentimes so pressed by other duties that little time remains for the preparation of experiments. Selected from the standard articles in the outfit of the chemist, they are neat in appearance, efficient in action, easily put together, and comparatively cheap. The forms shown in the cut are of the pieces which have been actually used in devising and testing the experiments described.

Fragile articles, such as test-tubes and flasks, should be bought in quantity, to allow for breakage. One balance will serve several workers. The same is true of the thermometer, where economy must be practiced, and a single graduated cylinder for each will do very well, although in that case all cannot, at one time, make Ex. 37.

The pieces are here described in the order of their numbers in the cut.

**1. Graduated Cylinder.** — A tall and narrow glass cylinder on foot, about $\frac{5}{8}$ inches in diameter, 25 cc. graduated to halves.

Number required in single set . . . . . 2

**2. Test-tubes and Rack.** — Tubes 6 inches long by $\frac{5}{8}$ inch in diameter. They should be bought by the gross. The rack to support the tubes can be made by any carpenter. Its form is shown in the cuts.

Number required in single set . . . . . 12

**3. Side-neck tube.** — Of German or soft glass, and of Bohemian or hard glass, 6 × ⅝ inches. They may be bought by the dozen, and kept in stock.

| | |
|---|---|
| Number required in single set, of soft glass . . | 1 |
| " " " " hard " . . . | 1 |

**4. Mortar and Pestle.** — Of glass or porcelain, glazed inside and outside, 3½ inches diameter.

**5. Support with Ring and Clamp.** — The support is of iron. It is the so-called retort-stand. We select the small size, usually provided with two rings. Only one ring is called for, but the second will be found useful. These rings should be arranged, as shown, to be taken off at the side of the rod. The clamp is "small size" with "universal movement," to hold a tube or flask in horizontal or vertical or oblique position.

**6. Side-neck Flask.** — A round bottom flask with a side neck attached to the stem.

| | |
|---|---|
| Number required for each student's set, 150 cc. . | 1 |
| " " " the teacher's set, 150 cc. . | 1 |
| and 250 cc. . | 1 |

**7. Conical Flasks.** — The so-called Erlenmeyer flasks, or Beaker flasks. These are thin and rather fragile, but with care will last a long time. Their freedom from color, perfect transparency, uniformity in shape and size, render them peculiarly well fitted for the examination of gases and liquids. Flasks of the usual form, or bottles, may be used instead of these, and, if made of white glass, and have mouths of uniform size, to be perfectly closed by the stoppers, bottles are very good and more durable. The mouth should be ¾ to 1 inch in diameter, taking a No. 4 or No. 5 E. & A. soft rubber stopper.

The use of these flasks and bottles, as described, for collecting and examining gases, dispenses with the pneumatic cistern, and the unpleasant wetness which goes with it. This method also enables one to dispose of noxious substances with gratifying success.

| | |
|---|---|
| For each student 200 cc. or 8 oz. flasks . . . . . | 4 |
| For the teacher, 200 cc. or 8 oz. flasks . . . . . | 2 |
| 400 cc. or 16 oz. flasks . . . . | 4 |

**8, 9. Rubber stoppers.** — Success in managing gases by this method demands that all joints in the apparatus shall be *air-tight.* Such joints are easily made by means of rubber stoppers No. 9, arranged with glass tubes No. 8. Each flask or bottle is to be supplied with this arrangement, and they are then to be joined together by rubber tubing. The first cost of rubber stoppers is larger than of cork, but their durability is a compensation. They should be of the "best soft rubber," and of such sizes as to fit the mouths of the flasks and tubes in use. Their sizes are described by numbers, but the same number, used by different makers, does not always describe the same size or quality. To be definite, we now refer to those stamped "E. & A." (Eimer and Amend).

No. 3 will fit a $\frac{5}{8}$ inch tube or flask or bottle.
No. 4 " " $\frac{3}{4}$ " flask or bottle.
No. 5 " " 1 " " bottle.

One "solid" will be needed, to close the side-neck tube or flask; one "with two holes," to close each flask or bottle in use.

**10. Bunsen Burner.** — In laboratories not supplied with gas, the alcohol lamp is the best substitute.

**11. Glass Funnel.** — The best German make, with long thin stem. Diameter $2\frac{1}{2}$ inches.

**12. Wide-mouth Bottles.** — Flint glass, tall style.

Bottles should be bought by the dozen, or in larger laboratories by the gross.

| | |
|---|---|
| Needed for each set, wide-mouth, 200 cc. or 8 oz. . | 1 |
| extra wide-mouth, 200 cc. " " . | 1 |
| wide-mouth, 400 cc. " 16 oz. | 1 |

**13. Forceps.** — Steel, plain, $4\frac{1}{2}$ inches.

**14. Glass tubing.** — Made of the best German or soft glass. It should be of such size as to fit the holes in the rubber stoppers used, about $\frac{3}{16}$ inch outside diameter for those described above. It may be cut into pieces, of any length desired, by first drawing across it the edge of a sharp three-cornered

file, once, making a distinct scratch, and then pulling the tube, almost but not quite lengthwise. The sharp cut edges may be rounded by heating them until red in the lamp-flame.

Glass tubing is to be bought by the pound.

**15. Porcelain Dish.** — So-called "evaporating dish." The R. B. porcelain is best. Diameter 3½ inches.

These may be bought by the dozen.

Needed for each set . . . . . 1 or 2

**16. Drying Tube.** — The so-called "chloride of calcium tube," with one bulb, length 5 inches.

**17. Pinch-cock.** — Mohr's, small size, strong.

**18. The Balance.** — A balance of good quality is the most costly piece in the outfit for laboratory work. In qualitative chemistry, such as the foregoing course, it is not absolutely indispensable, because something can be done by means of cheap substitutes, — even by such as an ingenious student can make.

The balance shown in the cut is a Becker's balance, listed in "Becker Brothers" catalogue as No. 14, at $11.00. With a glass case, which is very desirable to protect the instrument from dirt and corrosion, it is listed as No. 16, at $22.00. It is neat, accurate, sensitive to 2 mg., and durable.

The weights for this balance should be a set of 50 g. to 1 g. in brass, and 500 mg. to 1 mg. in platinum or aluminum with forceps, all in a lined and covered box. Such a set is listed by Becker Brothers at $9.00, and by Eimer and Amend at $5.50.

From parties (John Wannamaker, Philadelphia) who import from Becker's Sons of Rotterdam, the balance and weights just described, or their equivalent, can be obtained by schools, free of duty, at much less cost.

**19. The Water Pan.** — A pan about 8 inches diameter, and 3 or 4 inches deep, with flat bottom and straight walls. It may be of glass — a so-called crystallizing dish, as shown in the cut, or it may be of agate-iron ware, which is likely to be more durable, but less shapely.

**20. A common Plate.** — A small size china plate.

**21. Rubber Tubing.** — To connect the Bunsen Burner with the gas supply, white rubber, thick, $\frac{1}{4}$ inch diameter inside, may be used. Red rubber is better, and a little more costly.

For joining parts of the gas-apparatus the black or red tubing, of usual thickness, is to be preferred. The size should correspond with that of the glass tubing, which it must fit. Rubber tubing is bought by the foot. In pieces of 12 feet it comes a little cheaper.

**22. A Chemical Thermometer.** — A Centigrade thermometer, graduated from about —20° to +200°. The best instrument has its scale on the glass stem itself. A cheaper and very good instrument has a paper scale enclosed in a glass tube, which protects the stem.

## THE CHEMICALS.

In the following list may be found the names and formulas of all the substances required to make the experiments described in this book. Chemicals, to be used in the study of chemistry, should be of the best quality. Many of those furnished by the shops are impure, and often lead to wrong and troublesome results. It is better to buy chemicals, as you buy apparatus, from well-known dealers in laboratory supplies.

Reagents, which are to be used by students, should be kept upon their tables in small bottles: liquids in glass-stoppered bottles holding about 125 cc. or four oz., and solids in salt-mouth bottles holding 2 oz. If substances are to be used by the teacher they may, for the most part, be kept in the bottles in which they are bought. Every bottle should be distinctly and permanently labelled.

Unless economy must be rigidly practiced, the supply will not be limited to the substances in this list; specimens, in great variety, are very desirable.

The author will gladly give any information he can in regard to the purchase or use of apparatus and chemicals.

| | |
|---|---|
| Acetic acid, pure | $HC_2H_3O_2$ |
| Alcohol | $C_2H_6O$ |
| Alum | $K_2Al_2(SO_4)_4 + 24H_2O$ |
| Ammonium carbonate, C. P. | $(NH_4)_2CO_3$ |
| " chloride, C. P. | $NH_4Cl$ |
| " hydrate | $NH_4HO$ |
| " nitrate, cryst. | $NH_4NO_3$ |
| Antimony chloride, sol. C. P. | $SbCl_3$ |
| Arsenous oxide | $As_2O_3$ |
| Barium chloride, C. P. | $BaCl_2$ |

| | |
|---|---|
| Bismuth nitrate, cryst. C. P. | $Bi(NO_3)_3$ |
| Bone-black | C |
| Bromine | Br |
| Calcium chloride, crude | $CaCl_2$ |
| " " cryst. C. P. | |
| " oxide (quicklime) | CaO |
| Carbon pencil | C |
| Chrome alum | $K_2Cr_2(SO_4)_4 + 24H_2O$ |
| Cobalt nitrate, C. P. | $Co(NO_3)_2$ |
| Cochineal | |
| Copper, thin sheet | Cu |
| " chloride | $CuCl_2$ |
| " sulphate, C. P. | $CuSO_4$ |
| Dutch metal (imitation gold-leaf) | Cu Zn |
| Ferrous sulphate, pure | $FeSO_4$ |
| " sulphide (sticks) | FeS |
| Hydrochloric acid, pure | HCl |
| Iodine | I |
| Lead acetate | $Pb(C_2H_3O_2)_2$ |
| Litmus (blocks) | |
| Logwood | |
| Magnesium (ribbon) | Mg |
| " chloride, cryst. | $MgCl_2$ |
| " sulphate, C. P. | $MgSO_4$ |
| Manganese dioxide, powder | $MnO_2$ |
| " sulphate | $MnSO_4$ |
| Marble | $CaCO_3$ |
| Mercury, redistilled | Hg |
| Mercuric chloride (cor. sub.) | $HgCl_2$ |
| " oxide | HgO |
| Nickel chloride | $NiCl_2$ |
| Nitric acid, pure | $HNO_3$ |
| Oxalic acid, C. P. | $H_2C_2O_4$ |
| Paraffine | $C_nH_{2n+2}$ |
| Phosphorus | P |
| Potassium | K |
| " bromide, C. P. | KBr |
| " carbonate, C. P. | $K_2CO_3$ |

| | |
|---|---|
| Potassium chlorate, cryst. | $K Cl O_3$ |
| " chromate | $K_2 Cr O_4$ |
| " dichromate | $K_2 Cr_2 O_7$ |
| " ferrocyanide | $K_4 Fe Cy_6$ |
| " hydrate, pure | $K H O$ |
| " iodide, C. P. | $K I$ |
| " nitrate, cryst. | $K N O_3$ |
| " sulphate, cryst. | $K_2 S O_4$ |
| Platinum foil | Pt |
| " wire | Pt |
| Pyrogallic acid | $C_6 H_3 (H O)_3$ |
| Silver nitrate, cryst. | $Ag N O_3$ |
| Sodium | Na |
| " biborate (borax) | $Na_2 B_4 O_7 + 10 H_2 O$ |
| " carbonate, C. P. | $Na_2 C O_3$ |
| " hydrate, C. P. | $Na H O$ |
| " nitrate | $Na N O_3$ |
| " sulphate | $Na_2 S O_4$ |
| Strontium chloride | $Sr Cl_2$ |
| " nitrate, C. P. | $Sr (N O_3)_2$ |
| Sulphur, flowers | S |
| " roll | S |
| Sulphuric acid, pure | $H_2 S O_4$ |
| Tartaric acid, cryst. | $C_4 H_6 O_6$ |
| Tin (granulated) | Sn |
| Zinc (sheet) | Zn |
| " (granulated) | Zn |
| Sugar (granulated) | $C_{12} H_{22} O_{11}$ |
| Salt | Na Cl |
| Charcoal | C |

## METRIC AND ENGLISH MEASURES.

### Measures of Weight.

| | | |
|---|---|---|
| 10 milligrams, mg. | = 1 centigram, cg. | = 0.154 grains. |
| 10 centigrams, cg. | = 1 decigram, dg. | = 1.543 grains. |
| 10 decigrams, dg. | = 1 GRAM, g. | = 15.432 grains. |
| 10 grams | = 1 decagram | = 154.323 grains. |
| 10 decagrams | = 1 hectogram | = 3.527 oz. avoir. |
| 10 hectograms | = 1 kilogram | = 2.204 lb. avoir. |

1 grain = 0.0648 g. or 64.799 mg.
1 oz. Troy = 31.1035 g. 1 oz. avoir. = 28.349 g.
1 lb. Troy = 5760 grains. 1 lb. avoir. = 7000 grains.

### Measures of Volume.

10 cubic centimeters, cc. = 1 centiliter, cl. = 0.338 fld. oz.
10 centiliters = 1 deciliter = 0.845 gill.
10 deciliters = 1 LITER, l. = 1.057 quart.
10 liters = 1 decaliter = 2.642 gal.
10 decaliters = 1 hectoliters, hl. = 26.417 gal.
10 hectoliters = 1 kiloliter = 264.18 gal.

1 cu. in. = 16.386 cc. 1 U. S. quart = 0.9469 l.
1 liter = 61.027 cu. in. 1 U. S. gal. = 3.785 l.
1 U. S. gal. = 231 cu. in. 1 Imp. gal. = 277.25 cu. in.

### Measures of Length.

10 millimeters, mm. = 1 centimeter, cm. = 0.3937 in.
10 centimeters = 1 decimeter = 3.937 in.
10 decimeters = 1 METER, m. = 39.37 in.
10 meters = 1 decameter = 32.8 ft.
10 decameters = 1 hectometer = 328.08 ft.
10 hectometers = 1 kilometer, km. = 0.62137 mile.

1 inch = 2.534 cm. 1 millimeter = 0.0393 in.
1 yard = 0.9144 m. 1 meter = 1.0936 yd.
1 mile = 1.6093 km. 1 kilometer = about $\frac{5}{8}$ mile.

### Measures of Temperature.

Freezing-point of water = 0° Centigrade, C. or 32° Fahrenheit, F.
Boiling-point of water = 100° " or 212° "

1° C. = $\frac{9}{5}$°, or 1.8° F. 1° F. = $\frac{5}{9}$°, or 0.555° C.

To change a Centigrade temperature to its equivalent Fahrenheit temperature: Multiply by $\frac{9}{5}$ and add 32° to the product.

To change a Fahrenheit temperature to its equivalent Centigrade temperature: Subtract 32 and multiply the remainder by $\frac{5}{9}$.

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